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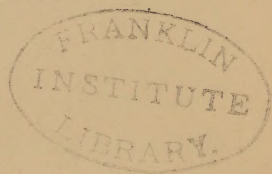offender shall be made public.
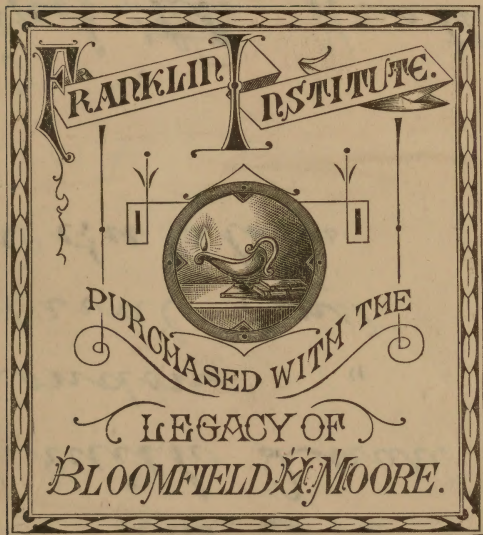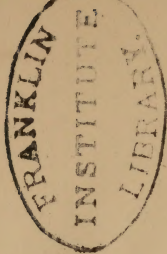
ELEMENTS

OF

CHEMISTRY.

BY

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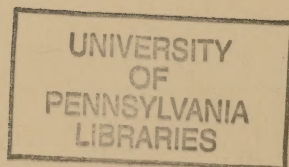MEMBER OF THE SOCIETY OF PUBLIC ANALYSTS OF ENGLAND.
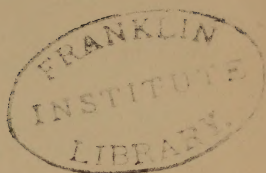
PREFACE.

THE first fifty-two pages of the present work comprehend a concise summary of chemical principles arranged for progressive study. The remainder of the work is an enumeration of the principal facts of the science, sufficiently elaborate for those attending college lectures or pursuing the common-school course of study. The experience of a dozen years, during which I have instructed over two thousand pupils, has convinced me that the most satisfactory progress is attained by a preliminary instruction in the principles of the science before proceeding to its descriptive part. It is on this method that this work is arranged. Teachers, however, who prefer to begin at once with descriptive chemistry, and to introduce the discussion of chemical principles from time to time in the course, will find the subject-matter so classified and indexed as to allow such system to be followed without difficulty.

The arrangement of the descriptions has been adopted as that best suited to render the acquirement of the facts easy, since bodies most alike in chemical relations are described in immediate succession. It has not been deemed necessary to follow the practice of the older manuals and defer the description of each compound body until all its elements have been described; in this book important com-

pounds have generally been presented in connection with their characteristic elements.

The science of Chemistry has become very difficult in the last twenty-five years, not only from the immense number of facts which have been presented, but also from the complexity of the theory, notation and nomenclature. That this manual may serve to lighten the labors of those who are endeavoring to acquire a knowledge of one of the most interesting and valuable of sciences, is the hope of

THE AUTHOR.

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ELEMENTS OF CHEMISTRY.

1. INTRODUCTORY.

CHEMISTRY is the science which investigates the composition of bodies and the changes which occur in them under various influences. Of the many forms of change which occur in nature, Chemistry studies especially those which produce some alteration of composition, though of late years many actions have been noticed in which the only change appears to be in the *arrangement* of the particles of the bodies, and many such changes are considered chemical phenomena.

Material substances may exist in at least three forms—solid, liquid and gaseous. The passing of any body from one of these forms to the other is generally brought about by heating or cooling, and is called a *physical*, not a chemical, effect. In the same way, the condition called magnetism is not known to be attended by any alteration of composition. On the other hand, the rusting of iron, the combustion of coal, the decay of animals or plants, give rise to new substances, and are strictly within the province of chemistry.

The material objects around us present much variety in color, form and general qualities. We can, by mechanical action, reduce most of them to fragments or even to fine

powder. Chemistry has taught us that many bodies are capable of another kind of division, which consists in separating them into two or more substances unlike each other and unlike the original substance. This change is called DECOMPOSITION. A fragment of common salt, for instance, may be converted into a fine powder, but by other means it may be decomposed into two substances—one a brilliant solid, and the other a greenish-colored gas. These two substances seem to be incapable of further decomposition, and are called ELEMENTS. Chemists, by experimenting upon a great number of bodies, have established the existence of sixty-six elements, to each one of which a name has been given. A certain force, or form of energy, known as CHEMICAL AFFINITY, causes these elements to unite with one another and form compounds. Some of the elements occur in nature in the free or uncombined condition, but most of them are in combination; hence the various objects of chemical study are either *elementary* or *compound*. The elements are usually arranged in two groups or classes—METALS and NON-METALS; but such division is unsatisfactory. A better arrangement is into three groups—NON-METALS, METALLOIDS and METALS. Even this, however, is not of much service, and will not be strictly adhered to in this work. The following arrangement into small groups is better for the purposes of study, and should be memorized sufficiently to enable the student to give without hesitation the members of the different groups, as this knowledge will be of great assistance in the study of the important compounds:

Oxygen Group.	*Chlorine Group.*	*Nitrogen Group.*	*Carbon Group.*
Oxygen,	Chlorine,	Nitrogen,	Carbon,
Sulphur,	Bromine,	Phosphorus,	Silicon,
Selenium,	Iodine,	Boron,	Tin,
Tellurium.	Fluorine.	Arsenic,	Platinum.
		Antimony,	
		Bismuth,	
		Gold.	

Potassium Group.	*Calcium Group.*	*Zinc Group.*	*Iron Group.*
Potassium,	Calcium,	Zinc,	Iron,
Sodium,	Barium,	Magnesium,	Manganese,
Lithium,	Strontium,	Cadmium.	Aluminum,
Hydrogen,	Lead.		Chromium,
Silver.			Nickel,
			Cobalt.

Copper Group.
Copper,
Mercury.
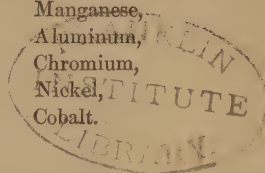
More complete lists are given in the latter part of the book.

Atomic Theory. If we reduce a piece of sulphur or any other element to powder, it will be seen to be capable of being further divided, and it would seem as if no limit existed to this division. Chemists, however, are now generally of the opinion that a limit does exist, and that every substance is made up of particles of definite size incapable of diminution or destruction. These have never been demonstrated. The belief in their existence is dependent upon some difficult mathematical and physical considerations; they are infinitely small, and equally hard, no matter what the nature of the mass which they make up. These particles are called ATOMS (a word derived from the Greek, and signifying indivisible), and any mass of elementary matter consists of a collection of a greater or less number of these atoms. It is now a generally accepted theory that even in elementary bodies the atoms are not perfectly free, but associated in pairs. A combination of atoms is called a MOLECULE. When, therefore, we divide a piece of sulphur, we merely separate the molecules from each other. Sulphur is soft, carbon in the form of the diamond is very hard, but this difference in hardness between the two is due simply to the firmness with which the molecules of the two bodies hold together. Those of the diamond have a strong affinity, those of sulphur a weak affinity; but in each body the molecules themselves are composed of atoms of infinite hardness.

When a solid becomes a liquid or a liquid becomes a gas,

or the reverse occurs, the molecules are not changed, but merely separated from one another. Hence the atoms in sulphur vapor are as hard and solid as those of solid sulphur, but in the vapor the pairs or molecules which they form are separated by greater distances than in the case of the solid. The following will render this point clear:

aa aa aa aa			The element in the solid state.
aa aa aa aa		"	" liquid "
aa aa aa aa		"	" gaseous "

This does not represent the proportionate separation, but only the general idea that this change of state is a simple separation of molecules. According to this view, bodies in passing from solid to liquid, and from liquid to gas, should increase in volume, and in the reverse processes should decrease; and this we find to be the case with the vast majority of substances.

The force which holds *atoms* together and forms them into molecules is a chemical force, and is called CHEMICAL AFFINITY. Any number of *molecules* of the same kind may be held together in a mass; the force that does this is called COHESION. Molecules of different kinds may also be held together; such an effect is called ADHESION.

Atomic Weights. Chemists have never been able to isolate or render visible atoms or molecules. Their size and weight remain entirely a matter of speculation and theory. Nevertheless, the progress of chemical research and the application of mathematics have developed some general principles. These are:

1st. That the atoms of each element have a constant and definite weight.

2d. That the atom of hydrogen is the lightest of all.

3d. That combination takes place among atoms under the action of chemical affinity.

Starting with the first two principles, numbers have been obtained which represent the weight of each atom compared to the atom of hydrogen. These numbers are called ATOMIC WEIGHTS. A complete list is given at the end of the book. It is not necessary for the student to commit them to memory.

2. NOTATION.

A CHEMICAL symbol is an abbreviation of the name of an element; in most cases an initial letter is used, as C for carbon, P for phosphorus. As some elements have names beginning with the same letter, proper distinction is obtained by assigning the single letter to the most common, and attaching small letters to the other initials. Thus, C stands for carbon, Ca for Calcium, Cl for chlorine, Cd for cadmium. Certain elements have different names in different languages, and for these the symbol is formed from the Latin name. Iron, for instance, is represented by Fe (*ferrum*); lead by Pb (*plumbum*); silver by Ag (*argentum*); potassium by K (*kalium*).

Symbols are absolutely invariable. No symbol represents two elements, no element has more than one symbol. The student should commit to memory thoroughly and accurately the symbols of all the important elements. A complete list of them will be found with the table of atomic weights.

To express combination between elements—in other words, to express the composition of a compound body or the molecules of an elementary body—the symbols are to be written together like the letters of a word. Such a collection of symbols is called a FORMULA.

The symbol, however, not only represents the element, but also one indivisible particle of it; that is, one atom. Hence, the expression CaO not only shows a compound consisting of calcium and oxygen, but also indicates that it contains a single

atom of each element. CaO_2 shows that two atoms of oxygen are present and one of calcium. In writing these expressions certain rules are followed:
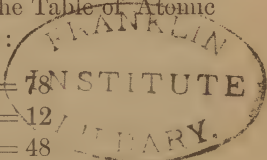
1st. To multiply any single atom, a small number is attached to the lower right hand, as seen above, where O_2 indicates two of oxygen. The formula $C_2H_4O_2$ shows a combination consisting of two atoms of carbon, four of hydrogen and two of oxygen.

2d. To multiply several atoms by the same number, we put a large figure in front. Thus $2HClO$ is equal to $H_2Cl_2O_2$; that is, the large figure multiplies the whole expression. This rule gives much trouble to beginners.

3d. To multiply a portion of an expression, several methods are in use. We may enclose the part to be multiplied in parentheses, and attach the proper number to the lower right-hand corner. $Ba(NO_3)_2$, for instance, equals BaN_2O_6; $C_6H_8(NO_2)_2O_5$ equals $C_6H_8(N_2O_4)O_5$. The effect of the small figure is limited to the part within the parentheses. This method is especially adapted to multiplying symbols in the middle or at the end of a formula. To multiply the symbols at the beginning of a formula, we usually point off or punctuate the part to be affected, and place a large figure in front. Some irregularity prevails as to the particular sign used, the comma and semicolon both being employed. It is sufficient for the student to bear in mind that a punctuation-mark or plus-sign occurring in a formula will stop the multiplying effect of the large figure at the beginning of the expression. For instance, $2C_2H_5,H_2N$ is equal to $C_4H_{10}H_2N$; similarly, in $2FeSO_4 + HCl$ the letters following the plus-sign are not affected by the figure 2. If we wish to carry the multiplying effect to the end of the expression, we enclose it in parentheses; thus, $2(FeSO_4 + HCl)$. Here all the letters are equally influenced.

Since the symbol of each element represents one atom, it follows that every symbol carries with it an idea of quantity.

If we write HCl, the meaning is not merely that hydrogen and chlorine are in combination, but that the amounts by weight are in the proportion of the atomic weights; *i. e.* 1 (atomic weight H) to 35.4 (atomic weight Cl). When the symbol is multiplied, the weight is also multiplied. For instance, H_2O represents 2 parts by weight of H to 16 of O; $HgCl_2$ represents 200 parts of mercury and 70.8 (35.4×2) parts of chlorine. From these examples it appears that formulæ give no exact idea of the percentage of the different elements unless we multiply each symbol by its atomic weight. The following illustration will perhaps make this point clearer: By burning sulphur in air or in oxygen we obtain a gas which contains equal parts by weight of S and O. To use the formula SO would be incorrect, for the table of atomic weight shows that the atom of S equals 32, while that of O equals 16; SO, therefore, would show a relation of 32 to 16. The relation is 32 to 32—that is, equal parts; hence, we must use the formula SO_2, which gives us $S = 32$; $O_2 = 32$ (16×2), the proper proportion. In the same way we may reduce a more complicated formula. Potassium carbonate is written K_2CO_3. Referring to the Table of Atomic Weights, we find the following numbers:

$$K = 39, \text{ consequently } K_2 = 78$$
$$C = 12, \qquad\text{``}\qquad C = 12$$
$$O = 16, \qquad\text{``}\qquad O_3 = 48$$

The sum of the atomic weights is called the MOLECULAR WEIGHT. In the example above given the sum is 138; we cannot reduce this sum in any way except by taking away one of the atoms; for, by the atomic theory, we cannot remove a portion of an atom. To take away any atom is to change the composition of the body; it is no longer potassium carbonate. Therefore, we say that the smallest portion of potassium carbonate that can exist will be 138 times as heavy as one atom of hydrogen. The same is true of the molecular weight of any body. It will represent the relation

2

between the weight of the smallest possible quantity of the body, its molecule, and one atom of hydrogen. With a number of compound gases it has been found that the molecular weight is equal to twice the specific gravity compared to hydrogen. The inferences from this fact will be discussed later.

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3. NOMENCLATURE.

THE names of chemical compounds are regulated by a system which depends essentially upon the employment of certain terminations.

In the old division of the elements into metals and non-metals the metals were usually distinguished by the termination " UM." A change of this termination into " A " indicated combination with oxygen. Potassium (K) becomes by oxidation potassa (K_2O); sodium (Na) becomes soda (Na_2O); magnesium (Mg) becomes magnesia (MgO). As the names of many of the common metals did not end in " um " unless the objectionable Latin name was used, this rule was only of limited application. The tendency of the modern nomenclature is to make but little change in the names of the substances called metals, and the terminations about to be presented are not usually attached to bodies ending in " um," or to those which we commonly call metals, such as iron, silver and zinc.

Chemical compounds which contain only two elements are called BINARY COMPOUNDS. They are usually named by joining the names of the elements present and attaching to one of them the termination " IDE." This termination may be conveniently regarded as an equivalent of the phrase " nothing else;" that is, wherever it occurs it indicates that nothing else is present except what is expressly mentioned. Potassium iod*ide*, for instance, can contain nothing else but

potassium and iodine; copper sulph*ide* can contain nothing but copper and sulphur.

PbO	Lead ox*ide*.
NaCl	Sodium chlor*ide*.
AgBr	Silver brom*ide*.

The syllable "ide" is usually attached to the members of the oxygen, chlorine, nitrogen and carbon groups, and preferably to those of the first two groups. Thus, a compound of iron and carbon is called iron carbide, but a compound of carbon and chlorine is called carbon chloride.

In many books, especially in older works, the word "of" will be found frequently used in the names of compounds. Instead of copper sulphide, we may see sulphide of copper, iodide of potassium for potassium iodide. This system was introduced into chemistry by an original mistranslation of French phrases in which the word "*de*" occurred. Good usage has fortunately justified the entire omission of the word in English chemical nomenclature.

As elements may combine in several proportions, forming several different compounds, this termination *ide* does not suffice for proper distinction, and chemists employ a system of prefixes used with this termination. These prefixes are formed mostly from the Greek or Latin numerals. The bodies Cu_2O and CuO are both properly called copper oxide, because they contain only copper and oxygen, but they are quite different substances. In the same way, SO_2 and SO_3 are both sulphur oxides, but must be distinguished from each other. The distinction is made in this way:

Cu_2O	Copper *sub*oxide.
CuO	" *mon*oxide (also *prot*oxide).
SO_2	Sulphur *di*oxide (also *deut-* or *bin*oxide).
SO_3	" *tri*oxide (also *ter*oxide).

The terms in parentheses are now rarely used. In regard to the first prefix, *sub*, it is of rather uncertain meaning. It

generally indicates deficiency; that is, that the quantity of the element to which it is attached is less than it should be. We apply the term *sub* especially to compounds in which a member of the oxygen or chlorine group is deficient in amount. Pb_2Cl, Zn_3I_2, Cu_4O are sub-compounds.

With some of the elements, however, the *normal* proportion of combination is two of the first to one of the second, and here it has become universal custom to use the prefix *mon*. We have, therefore,

K_2O Potassium *mon*oxide,
Na_2O Sodium *mon*oxide,
Ag_2O Silver *mon*oxide,
K_2S Potassium *mon*osulphide,

and so on with all the series. The following includes the elements which show this exception:

Chlorine,   Hydrogen,
Bromine,   Potassium,
Iodine,   Sodium,
                     Lithium,
                     Silver.

It is especially in the combinations of these bodies with members of the oxygen group that the irregularity is shown.

Some elements form compounds in which the proportion is as 1 to $1\frac{1}{2}$, but as fractions are not allowed in formulæ, the whole expression is multiplied by 2, which gives the proportion 2 to 3. $FeO_{1\frac{1}{2}}$ becomes, therefore, Fe_2O_3. These are called *sesqui*-compounds, and the above expression is iron sesquioxide. The word *sesqui* means one and a half, and conveys the idea that the relation between the two elements is as 1 to $1\frac{1}{2}$ (2 to 3). Higher proportions also occur. We have, for instance,

CCl_4 Carbon tetrachloride or quadrichloride.
PCl_5 Phosphorus pentachloride.

In assigning names to compounds containing more than two elements, a great difficulty occurs from the very large number of such compounds which may be formed. To express the names of all the elements would often make long words, so the general custom is either to leave some elements unexpressed or to give a name to a group of two or more of the elements. For instance, the substance KHO is called potassium hydrate. In this name only the K and H are mentioned; the O is indicated by the termination. The compound KCN is called potassium cyanide, the combination CN being called cyanogen. This latter method of giving special names to groups is very common in organic chemistry.

Among the most important of the compounds containing three elements are those which we call SALTS. This term is difficult to define; it comes to us from the early days of chemistry, when opinions as to the nature of substances were quite different from those held at present. Briefly, salts may be defined as substances formed by the action of acids upon certain elements or their oxides. If we put zinc or zinc oxide into sulphuric acid, we get a zinc salt; in this case zinc sulphate. We can get salts also by direct union of many oxides; for instance, when calcium oxide, CaO, acts upon carbon dioxide, CO_2, we get calcium carbonate, $CaCO_3$, which is a salt.

Most salts contain *three* elements, of which *oxygen* is one, and the names are made by joining the names of the other two elements and adding to them certain syllables which not only indicate the presence of oxygen, but also partly the amount. These syllables are " ATE " and " ITE." The former indicates the greater quantity of oxygen. Thus potassium sulph*ate* and potassium sulph*ite* both contain oxygen, but the former (sulph*ate*) contains the more oxygen. Sodium nitr*ate* and sodium nitr*ite* contain the same elements, but their composition is $NaNO_3$ and $NaNO_2$, respectively.

It will aid in the comprehension of this subject if we extend

the principle which has already been mentioned when speaking of the termination "IDE." It was pointed out that this syllable could be regarded as equivalent to the phrase "nothing else." In the same manner, the syllables "ATE" and "ITE" are to be regarded as meaning "something else," and that something else is generally oxygen. With these points in mind the student will recognize, at a glance, that while in sodium sulph*ide* but two elements are present, sodium sulph*ate* and sulph*ite* will contain three.

These two terminations are not sufficient to distinguish all the salts that may be formed from certain elements. For instance, potassium, chlorine and oxygen will unite in four different proportions, forming $KClO_4$, $KClO_3$, $KClO_2$, $KClO$. In such cases the important or most common compound is distinguished by the termination "ATE," and the one containing the next lower amount of oxygen by the termination "ITE."

The other compounds are indicated by the use of certain extra syllables, "HYPO" and "HYPER," the latter now generally abbreviated to "per." The significance and use of these syllables are shown below:

$KClO_4$	Potassium perchlorate.
$KClO_3$	"　　chlorate.
$KClO_2$	"　　chlorite.
$KClO$	"　　hypochlorite.

From this table it is seen that "PER" intensifies the meaning of any termination—that is, indicates more oxygen than if the termination were used alone; while "HYPO" diminishes the power of a termination—that is, indicates a smaller amount of oxygen than would be present if "HYPO" were not used. Several other series of salts show the same principle, although not so perfectly as that above given:

Na_2SO_4	Sodium sulphate.
Na_2SO_3	"　　sulphite.
Na_2SO_2	"　　hyposulphite.

When, in such compounds, hydrogen is present, we might use a similar system, but the custom of chemists has decreed that a different method shall be adopted.

Taking the series of chlorine compounds given above, in place of the potassium salts we might have $HClO_4$, $HClO_3$, $HClO_2$, $HClO$, and these might be called hydrogen perchlorate, hydrogen chlorate, etc.; such names are used by a few persons, but have not become current. The more frequent method is to drop the word "hydrogen," change the termination ATE into IC, the termination ITE into OUS, and add the word ACID.

The series would therefore be,

$HClO_4$	Perchlor*ic* acid.
$HClO_3$	Chlor*ic* acid.
$HClO_2$	Chlor*ous* acid.
$HClO$	Hypochlor*ous* acid.

Note particularly that the *prefixes* are retained without change, and that the syllable IC is found whenever, in the metallic salt, the termination was ATE, and the syllable OUS is found when the name has come from a compound ending in ITE.

Under this rule we have the following transformations:

Potassium sulphate } corresponds to { Sulphuric acid
   K_2SO_4 }   { H_2SO_4

Potassium sulphite } " { Sulphurous acid
   K_2SO_3 }   { H_2SO_3

Potassium hyposulphite } " { Hyposulphurous acid
 K_2SO_2 }   { H_2SO_2

Strictly speaking, no necessity exists for this variation. The compounds containing hydrogen ought to be regarded as salts. They exhibit, however, some incidental properties which distinguish them from the rest of the salts. They have a sour taste, redden vegetable blues, and have, as a rule, a wider range of chemical action. They stand out as a group,

and from a very remote period have been called acids. The term is too well established in chemistry to be set aside.

Sometimes we have bodies in which the hydrogen is only partly replaced by another element, and these are intermediate between the acids and the salts. Thus, $KHSO_4$ is at once a potassium and a hydrogen compound. In such cases the name is a combination of both systems. The above compound, $KHSO_4$, is called acid potassium sulphate. Here the word *acid* calls attention to the existence of hydrogen, while the rest of the elements are indicated by the latter part of the expression. These acid salts are not unfrequently called *bi*-salts. Acid potassium sulphate, for instance, is generally known in commerce as potassium bisulphate; the corresponding acid carbonate, $KHCO_3$, as bicarbonate. This use of the syllable *bi* is improper. If it means anything in this connection, it is that two molecules of acid are present, which is not the case. In a few compounds of exceptional composition the title is used for want of a better one. $K_2CrO_4CrO_3$, for instance, is called potassium bichromate. It is not properly so called; it does not contain two molecules of chromic acid, for CrO_3 is chromium teroxide; but the more scientific title, anhydrochromate, will not be likely to find favor, and the incorrect name will long be used.

The terminations "ous" and "ic" were formerly used only for acids or for the direct derivatives from them. Of late years these syllables have been much employed for distinction in cases in which an element forms two sets of compounds. The following instances will explain this:

Mercury forms two chlorides, two iodides, two sulphides, etc. The two series are distinguished as follows:

Hg_2Cl_2	Hg_2I_2	Hg_2O	Hg_2SO_4
$HgCl_2$	HgI_2	HgO	$HgSO_4$
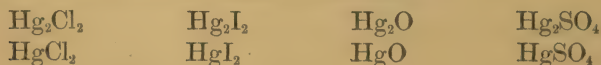
The bodies in the upper row are called mercur*ous* salts, those in the lower row mercur*ic* salts. In the same way we have

FeO	FeCl$_2$	FeSO$_4$	Ferro*us* salts.
Fe$_2$O$_3$	Fe$_2$Cl$_6$	Fe$_2$(SO)$_3$	Ferr*ic* "

Note particularly how these terminations are applied. They do not indicate the amount of the element to which they are attached, but of the other substance; *ous*, as usual, means less than *ic*.

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4. LAWS OF CHEMICAL COMBINATION.

CHEMISTRY is an exact science; that is, its laws and principles are mostly known and established. The great law of Chemistry is the LAW OF CONSTANT PROPORTION, and is expressed thus: *Every chemical compound is definite in its nature, containing always the same ingredients and in the same proportions.*

It must not be supposed from this that elements can only combine in one proportion; on the contrary, some of them, carbon and hydrogen for instance, combine in many proportions; but the important point conveyed in the above rule is that each of these compounds is a different substance. As an excellent illustration of this the compounds of mercury and chlorine may be mentioned. If we unite 200 grains of mercury with 35.4 grains of chlorine, we get calomel; if we unite 200 grains of mercury with 70.8 grains of chlorine, we get corrosive sublimate. The two bodies are so different that if we did not know their composition we would suppose them to contain different elements. Calomel is insoluble in water, and non-poisonous. Corrosive sublimate is soluble in water, and one of the most violent of poisons. When we examine a great variety of chemical compounds, we notice that often when elements combine in different proportions a simple arithmetical relation exists. A few cases will make this clear:

The two compounds of oxygen and carbon are known, CO and CO$_2$. Two compounds of sulphur and oxygen are known,

SO_2 and SO_3. In the series of compounds of potassium, chlorine and oxygen, given a page or two back, we have

$$KClO$$
$$KClO_2$$
$$KClO_3$$
$$KClO_4$$

in which the proportion of oxygen is seen to increase regularly. The observation of many facts of this kind has given rise to a sort of principle in chemistry which is called the LAW OF MULTIPLE PROPORTION; expressed thus: *When two bodies combine in more than one proportion, the higher proportions are multiples of the lower by a whole number.*

This law amounts to saying that if two elements are in combination in a low proportion, and we wish to add more of one of them, we must add one, two, three, four times, etc. as much. This principle, however, is not a law; it is true of the simpler bodies only; quite a number of elements combine in almost every proportion, and in many cases the multiple relation is only apparent, having been obtained by a sort of arithmetical trick. The great and important law of chemistry is that of constant proportion.

The atomic weights express the proportion in which the elements combine. The atomic weight of hydrogen, for instance, is 1; that of bromine 80; and the only known compounds of these elements contain 80 parts of bromine to one part of hydrogen.

In some elements the power of combination is such that more than one atomic weight of the one is required for one atomic weight of the other. Thus, the ratio of combination between H and O is H_2O; between N and H is H_3N; between C and H is H_4C. By substituting the atomic weights for the symbols we get the following proportions:

$H_2 = 2.$	$H_3 = 3.$	$H_4 = 4.$
$O = 16.$	$N = 14.$	$C = 12.$

If we reduce the above ratios to the simplest form, we find that

1 part by weight of H equals   8 parts of O.
1   "   "   H   "   4.7   "   " N.
1   "   "   H   "   3   "   " C.

These ratios are the simplest in which the elements can be compared, and are called EQUIVALENTS, because they represent the amount by weight in which the elements are equal to each other. They were formerly much used in chemical calculation and arrangement, but the atomic weight is now preferred. The equivalent of any two elements in reference to hydrogen will also be their equivalent with reference to each other. Thus, in the above examples oxygen is seen to be equivalent to hydrogen in the proportion of 8 to 1. Similarly, carbon is seen to be equal to hydrogen in the proportion of 3 to 1. We find, further, that carbon is equivalent to oxygen in the proportion of 3 to 8, for the compound they form is CO_2, in which, of course, $C = 12$ and $O = 32$, which is as 3 to 8.

Compound bodies combine with each other in the proportion of their molecular weights. Lime (CaO) absorbs carbon dioxide (CO_2), and forms calcium carbonate, $CaCO_3$. The proportion by weight in which they combine is determined thus :

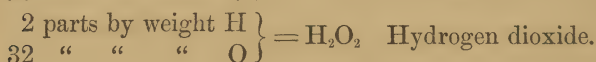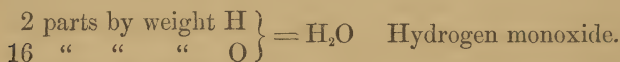Ca, atomic weight   40   C, atomic weight   12
O,   "   "   16 O_2,   "   "   32

CaO, molecular weight 56 CO_2, molecular weight 44

The proportion is therefore 56 parts of CaO to 44 parts of CO_2.

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5. ATOMICITY.

IT is mentioned above that elements may combine in several proportions, producing in each case distinct bodies. When

compounds containing the same elements are compared, we generally find one proportion which seems to be the most natural one; it is either most frequently or easily produced, or it is the one least liable to change. Hydrogen and oxygen combine in two proportions, thus:

$$\left.\begin{array}{l} 2 \text{ parts by weight H} \\ 16 \text{ " " " O} \end{array}\right\} = H_2O \quad \text{Hydrogen monoxide.}$$

$$\left.\begin{array}{l} 2 \text{ parts by weight H} \\ 32 \text{ " " " O} \end{array}\right\} = H_2O_2 \quad \text{Hydrogen dioxide.}$$

These two bodies are very different. The first is *water*. It is formed whenever hydrogen is allowed to burn in air, and it is well known as a compound not liable to decompose. The second substance is very difficult to prepare and to preserve; it is liable to explode. We can have no doubt, therefore, that the natural proportion of combination between H and O is H_2O. Carbon forms with oxygen two well-marked compounds, CO and CO_2. CO is formed when carbon is burned in a *deficient* supply of air, but CO_2 is formed when the carbon burns under natural conditions in a free draft of air or oxygen. CO, besides, shows a tendency to take up more oxygen, especially when heated, and it will combine with chlorine even at ordinary temperatures. CO_2, on the other hand, shows no tendency whatever to combine with oxygen or chlorine. This list might be continued at great length, for all the elements have been more or less extensively examined with reference to this point.

In developing this principle it has been found convenient to take the atom of hydrogen as a point of comparison, and to arrange each element according to the number of hydrogen atoms with which it forms the most permanent combination. Taking some important elements, for instance, we find their compounds with hydrogen as follows:

Cl combines with one H, forming HCl.
Br   "   "   " H,   "   HBr.
O   "   " two H,   " H_2O.
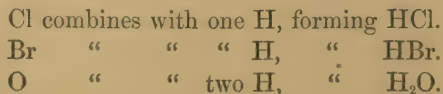
S combines with two H, forming H_2S.

N	"	" three H,	"	H_3N.
As	"	" " H,	"	H_3As.
C	"	" four H,	"	H_4C.

These are not the only compounds that can be formed from these elements, but they are those which show only a slight tendency either to take new atoms or give up what they already possess.

If an element does not form a compound with hydrogen, we may either combine it with some other body, and then compare that to hydrogen, or we may displace hydrogen from the combination, and thus get an idea of the number of hydrogen atoms to which the element is equal.

The greatest number of hydrogen atoms with which any element combines is called its ATOMICITY, QUANTIVALENCY, or VALENCY. Degrees of atomicity are indicated by names and signs; the signs are the Roman numerals, the names are derived from the Greek or Latin names for these numerals. One degree of atomicity is indicated by the mark ('), and the body so marked is called a *monad*, or is said to be monatomic or monivalent. In the same manner,

    " indicates a dyad (diatomic or bivalent).
    ‴   "   a triad (triatomic or trivalent).
    iv   "   a tetrad (tetratomic or quadrivalent).
    v   "   a pentad (pentatomic or pentivalent).
    vi   "   a hexad (hexatomic or hexivalent).

The terms ending in " *ad* " are most convenient for use.

Atomicity has nothing to do with the energy or activity of the element. It is a measure of capacity only. Bodies of high atomicity are often of weak affinity, while some of the strongest chemical agents are of low atomicity. Chlorine has only one-third the atomicity of nitrogen, but it is many times more energetic as a chemical substance.

Degrees of atomicity are determined by a study of the pro-

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portions in which bodies combine; hence a knowledge of the atomicity of the elements is a key to the composition of all their important and more permanent compounds. The following list should be thoroughly committed to memory, since it will enable the student to write correctly the formulæ of many common chemical substances:

The chlorine and potassium groups are monad.

The oxygen, calcium, zinc, copper and iron groups are dyad; the iron group is often hexad.

The nitrogen group is triad, and often pentad.

The carbon group is tetrad.

When elements are combined in such proportion that their atomicities are equal, the compounds are said to be SATURATED. This use of the word must be carefully distinguished from its older and more common use, meaning that a body has dissolved or absorbed as much of any substance as it can take up. In this sense we speak of *saturated solutions,* meaning solutions which contain as much of any substance as can be dissolved; we speak also of gases being saturated with moisture, meaning that they contain as much moisture as can be held by them under the conditions.

The above list can be used not only as a guide to the compounds which each element forms with hydrogen, but also as a guide to the compounds which the elements form with each other. The system is quite simple. Taking the monad group, for instance, it will be at once understood that as the members are all equal to one atom of H, they are equal to each other. Hence, K and Cl will combine in equal atoms, forming KCl, potassium chloride. Similarly we will have NaBr, AgI, etc. The dyad elements have twice the combining capacity of monads; we will find, therefore, that the compound of sodium and oxygen will be Na_2O.

The general rule is that the elements unite in such proportions that the degrees of atomicities are equalized. Suppose

we have a compound of Cl and Sb; Sb is a triad—that is, equal to three hydrogen atoms; while Cl is a monad, and equals only one hydrogen atom. It will therefore take 3Cl to have the same capacity as one Sb, and the proper formula will be $SbCl_3$. If a compound of C and Cl be desired, we find that as one is a tetrad and the other a monad, the resulting compound will be CCl_4.

The most difficult application of the rule is where triads and dyads are united, as in the compounds of oxygen and sulphur with nitrogen, phosphorus, arsenic and antimony. A few trials will show that to make the two atomicities equal we must multiply the dyad by 3 and the triad atom by 2. For example, Sb_2S_3 must be the formula of antimony sulphide, for

$$Sb_2 = 3 \times 2 = 6$$
$$S_3 = 2 \times 3 = 6$$

A few formulæ are here appended as additional illustrations:

Monad with monad			H'Cl'	K'Cl'	Na'Br'
"	"	dyad	H'_2O''		
"	"	triad	H'_3P'''	Ag'_3Sb'''	
"	"	tetrad	H'_4C^{iv}		
Dyad	"	dyad	$Cu''O''$	$Zn''S''$	$Fe''O''$
"	"	triad	$P'''_2O''_3$	$N'''_2O''_3$	$As'''_2S''_3$
"	"	tetrad	$C^{iv}O''_2$	$C^{iv}S''_2$	

It must, however, be borne in mind that compounds exist in which the proportion of the atoms differs from what the rule requires; but the essential character of such compounds is a tendency to change, either by taking new atoms or giving up some that they already possess. The normal compound of oxygen and carbon is, of course, CO_2, in which the one tetrad C atom is exactly saturated by the two atoms of dyad O. We are, however, acquainted with a well-defined body having the formula CO. This substance is rather a proof of the rule

than an exception to it, for it shows a strong tendency to take up other atoms in order to complete its structure. Heated in the air, it combines with oxygen and forms CO_2, and it unites with chlorine, forming a definite compound which will, of course, have the formula $COCl_2$, since it requires two chlorine atoms to perform the function of one oxygen atom.

So, also, the student must not fall into the error of supposing that bodies in which the degrees of atomicity are equalized are necessarily without chemical activity. On the contrary, some of our most active chemical agents are saturated compounds. Hydrochloric acid, HCl, is an example of such a substance. It shows no tendency to take new atoms except under the condition that it at first give up a portion of its structure. It will dissolve potassium, for instance, but only by first losing its hydrogen, into the place of which the potassium enters. It will dissolve potassa (potassium oxide, K_2O), but only by exchanging its hydrogen for the potassium. The changes may be represented as follows:

$$K' + H'Cl' = K'Cl' + H'$$
$$K'_2O'' + 2H'Cl' = 2K'Cl' + H'_2O''$$

This is, in fact, the nature of common chemical changes: they are substitutions of one element for another, the change always taking place in such a way that the element driven out is exactly equal in atomicity to the one that enters the combination. If, instead of acting on hydrochloric acid with potassium, we use zinc, the quantity of HCl will have to be increased; the reaction $Zn + HCl$ cannot take place, since one atom of zinc must drive out two atoms of hydrogen, zinc being a dyad. Therefore, we say $Zn + H_2Cl_2 = ZnCl_2 + H_2$, which is in strict accordance with fact, as showing that one atom of zinc will set free twice as much hydrogen as one atom of potassium will.

The degrees of atomicity given above are not invariable. The circumstances under which the variation takes place cannot be very well defined; but the extent or rate of varia-

tion is by a simple law, to which only a few exceptions need be made. *When an element changes its atomicity, either increasing or diminishing, the change is by two degrees at a time.* Elements of even atomicity remain even, passing, for instance, from hexads to tetrads, and finally to dyads, or *vice versâ;* elements of uneven atomicity remain uneven, passing from pentads to triads and monads.

Certain elements vary in atomicity in a way that appears to be exceptional, but in which we can, by a supposition, account for the change and yet preserve the application of the law. These bodies are supposed to have the property of combining with themselves in such a manner as to form double atoms, possessing an atomicity greater than either atom singly, but less than the sum of the atomicities of the two atoms. For instance, iron, which is generally a dyad, becomes in certain compounds a tetrad, but instead of forming compounds upon this basis, two atoms of iron unite and form a double atom or molecule, which then forms compounds with other elements. A short reflection will show that this molecule, formed from two atoms each having a capacity of four, will have a power of six, one degree of atomicity in each atom having been consumed in forming the compound.

For all cases of varying atomicity, whether regular or irregular, the terminations *ous* and *ic* are much employed, *ous* indicating the lower degree and *ic* the higher. We have in this way mercurous (lower atomicity) and mercuric (higher atomicity) salts, ferrous (dyad) and ferric (hexad) compounds. Indeed, in the use of the termination of the acids the same principle is carried out, sulphurous acid being the compound in which sulphur has a lower (tetrad) atomicity; sulphuric acid one in which sulphur has a higher (hexad) power.

In arranging formulæ containing three elements the application of the law of atomicity becomes somewhat difficult. In many common cases it will be found that the atomicity of one of the elements is much higher than it is in bodies

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containing two elements; sulphur, for instance, is a dyad in binary compounds, but in the sulphites and sulphates it is, respectively, a tetrad and hexad. When oxygen is one of the three elements, we usually count it against the sum of the other two. Taking an instance of the salts above mentioned, we would get the following formulæ:

$$Potassium\ sulphite \qquad K'_2 S^{iv} O''_3$$
$$Copper\ sulphate \qquad Cu'' S^{vi} O''_4$$

In each case the atomicity of the oxygen is equal to the sum of that of the other two elements.

For the student, however, the safest and shortest rule will be to commit to memory, thoroughly, some standard formulæ containing three elements, and from these, by very simple rules, a large number of compounds can be built up.

These formulæ are—

H_2SO_4	Sulphuric acid.
H_2SO_3	Sulphurous acid.
H_2CO_3	Carbonic acid.
HNO_3	Nitric acid.
HNO_2	Nitrous acid.
$HClO_3$	Chloric acid.
H_3PO_4	Phosphoric acid.

The derivatives from these bodies form a large part of common chemical substances. If we wish to write the formula of any metallic salt, we substitute the proper amount of metal for the hydrogen in the corresponding acid. Let it be required to write the formula of potassium carbonate; the reasoning would be as follows: Carbonic acid is H_2CO_3, potassium is a monad; two atoms of potassium will be required to substitute the two atoms of hydrogen, and the formula is K_2CO_3. By the same reasoning copper sulphate may be deduced. Sulphuric acid is H_2SO_4, copper is a dyad; one atom of copper will displace two of hydrogen; therefore, $CuSO_4$. When the standard formula contains too small an

amount of hydrogen, we must multiply the expression by some whole number. For instance, the formula of copper nitrate will be deduced in this manner; Nitric acid is HNO_3, copper is a dyad; copper will therefore replace the hydrogen of two molecules of nitric acid; hence, $Cu(NO_3)_2$ or CuN_2O_6.

Whenever we take one or more atoms from a saturated compound, we leave the compound unsaturated to a degree equal to the number of hydrogen atoms to which the removed atoms correspond. The molecule H_4C is saturated. The molecule H_3C is obtained by removing H, and is therefore a monad; H_2C is obtained by subtracting a second H, and is therefore a dyad; and so on. The atomicity of any molecule can thus be obtained by finding how much hydrogen is required to form a saturated compound. By this method we determine that HO is a monad, for it requires but one atom of H to complete the molecule; CO_3 is a dyad, for it requires H_2 to form the saturated compound H_2CO_3; PO_4 is a triad, for it forms H_3PO_4.

Graphic Formulæ. A convenient and much-used method of indicating atomicities is by GRAPHIC FORMULÆ. These consist of the symbol of each element, with bonds or prolongations the same in number as the degrees of atomicity. Taking some common elements as examples, we have

monad	dyad	triad	tetrad	pentad
K—	—O—	—P—	—C—	=N=
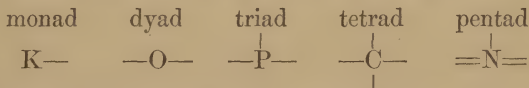
These bonds may be attached in any position or direction as long as the proper number is used. Carbon, for instance, may be written as above, or

$$C\equiv \qquad —C\equiv \qquad =C= \qquad \equiv C—$$

or in any other way, provided four bonds are present.

The practical application of this graphic notation to the writing of chemical formulæ is easy. We link together the bonds of the different elements, and when all the points are

joined the compound is complete and is a saturated molecule. Two bonds of one atom, however, can never be attached to a single bond of another atom. We cannot have $K—=O$, but $K—O—K$, showing us that the composition of potassium monoxide must be K_2O. The following are examples of some common compounds written graphically:

$$H—Cl; \qquad H—O—H; \qquad H—N{<}^{—H}_{—H};$$
$$O{=}C{=}O.$$
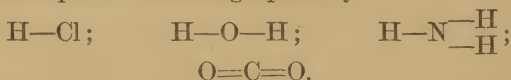
We may also indicate *unsaturated* molecules. Thus, $O{=}C{=}$ shows that carbon monoxide is a body having two degrees of atomicity unsatisfied; $O{=}C{=}^{Cl}_{Cl}$ that two atoms of chlorine have combined and satisfied this free atomicity.

The only objection, perhaps, to the use of graphic formulæ is the danger that the student may think that the atoms actually have spokes or projections on them, or are arranged in the somewhat architectural manner shown in the formulæ. $K—O—H$ does not mean that in potassium hydrate the atom of O is flanked on either side by a potassium and by a hydrogen atom, or that the atoms are connected by hooks or prongs, but merely that the oxygen atom has certain degrees of affinity which are satisfied by other atoms.

A special application of this notation is to explain the nature of those changes in atomicity which have already been mentioned. The atom of sulphur, for instance, is in some combinations a hexad, in others a tetrad, in others a dyad. This progressive diminution of capacity may be supposed to arise from the bonds of affinity combining with each other in pairs; thus:

Hexad S   Tetrad S   Dyad S
$=S=$ $=S=$ $(S=$
$\|$ $\smile$ $\smile$

The same principle can be shown with an element of uneven atomicity:

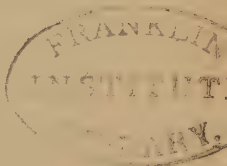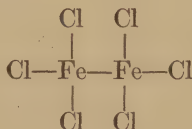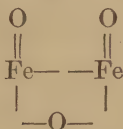| Pentad N	Triad N	Monad N
$\vert$	$\vert$	
$=N=$	$-N\,)$	$(\,N\,)$
	$\vert$	$\vert$

Since such combinations cannot take place unless both points are saturated or neutralized, the decrease of atomicity must take place by two degrees.

The nature of the change by which the iron atom passes from a dyad to a hexad condition can be very well shown by this method. Dyad iron, graphically represented, would be $(\,Fe=$, in which two bonds have satisfied each other, leaving two still active. In the higher degree of atomicity the condition is $-\overset{\vert}{Fe}-\overset{\vert}{Fe}-$, one bond of each atom having combined and linking the two in chemical union. Ferric oxide and ferric chloride would be

$$
\begin{array}{cc}
O \quad\quad O & Cl \quad Cl \\
\| \quad\quad \| & \vert \quad\ \vert \\
Fe-\ \ -Fe & Cl-Fe-Fe-Cl \\
\vert \quad\quad \vert & \vert \quad\ \vert \\
-O- & Cl \quad Cl
\end{array}
$$

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6. ELECTRICAL RELATIONS OF THE ELEMENTS.

ELECTRICAL excitement exhibits two opposite conditions, called respectively POSITIVE and NEGATIVE. These conditions are produced in any apparatus developing electricity. The points at which the electrical excitement is manifested—for instance, the wires of a battery—are called the POLES. The positive pole is usually distinguished by the sign $+$, and the negative by the sign $-$.

These have a strong tendency to unite and neutralize each other. On the other hand, positive electricity repels positive, and negative repels negative. Two bodies charged with dif-

ferent kinds of electricity will attract each other, but if charged with the same kind of electricity will repel. The law is generally expressed as follows: *Like electricities repel; unlike, attract.*

These principles have been applied to the determination of some important relations between elements. A current of electricity decomposes a large number of compound bodies, and some elements appear at the positive pole, and others at the negative. Thus, potassium will be liberated in contact with the surface negatively charged, and oxygen in contact with the positive surface. This will be the invariable result with these elements, no matter what compounds be taken for the experiment, but with many other elements the effect will depend upon the nature of the compound. With H_2S the sulphur will appear at the positive pole; with SO_2, at the negative. This difference is due to the superior attraction of oxygen to the positive pole; it seems to compel the sulphur to go to the other point.

Since unlike electricities attract, it follows that elements which go to the positive side must be negative, and those at the negative side must be positive. Very frequently we use the term " ELECTRO" in this connection; thus we say, zinc is *electro-positive;* chlorine is *electro-negative.*

A body is not absolutely positive or absolutely negative, but is simply more positive or more negative than some other substance. Nevertheless, as the list of elements is limited, we will have two bodies which, by their high affinities, will stand at the extremes of the scale, one being always negative, the other always positive. Leaving out of consideration some rare elements, we may place potassium as the most positive, oxygen as the most negative.

The following table shows the common elements arranged in the order of their electrical relations:

K, Na, Mg, Zn, Fe, Pb, Cu, Ag, Hg, H, C, P, S, I, Br, Cl, O.

Each element of this list will be positive when in combina-

tion with any element to the right of it, negative when in combination with any to the left.

These principles have an important application in determining chemical changes. The greatest chemical attraction generally exists between elements most widely separated in their electrical relations. We may, by this means, often discover the most probable result of any chemical action, as is shown in the next section.
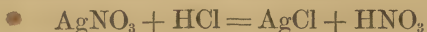
7. REACTIONS.

CHEMICAL symbols are employed not only to show the composition of bodies, but also to show exactly the nature of the chemical changes which occur when different bodies are brought in contact. When so used the expression is called a REACTION. Certain compounds, which are much used for producing reactions, are called REAGENTS, though strictly all the substances present take equal part in a reaction. When we pour vinegar upon a marble slab, we say, in ordinary phrase, that the marble is corroded, but, in fact, the vinegar is equally acted upon; both substances are changed in composition, both are rendered unfit for their original uses; in other words, they have not only acted; they have reacted, and are therefore both reagents.

A reaction is substantially an expression of the results of an experiment, and, when correctly written, gives us the proportion in which bodies are to be used and the proportion of the resulting substances. We can never be absolutely sure of the correctness of any reaction until we make the experiment and analyze the result; but the progress of Chemistry has made known certain laws of change which enable us to predict, or infer, many results without the necessity of actual observation. Every now and then, how-

ever, the analogy fails, and experiment disappoints the suggestions of theory.

Reactions are written by placing in proper proportion, connected by $+$ signs, the formulæ of the bodies concerned, then writing the sign $=$, and following this by the formulæ of the resulting bodies. For instance,

$$AgNO_3 + HCl = AgCl + HNO_3$$

expresses that on bringing together silver nitrate and hydrochloric acid a chemical change occurs by which silver chloride and nitric acid are produced. The principal difficulties in regard to reactions are: 1st. To know whether a given change will take place; 2d. To know the quantities of the bodies to be used; 3d. To know the nature of the resulting bodies. These points may be taken up in order.

1st. In the simplest cases the nature of the reaction will be determined by the affinities of the elements as governed by their electrical relations, the change taking place in such a way that the element having the stronger electric affinity will drive out and supplant the element of similar but weaker affinity. A reference to the table of electrical affinities will show that chlorine is more strongly negative than bromine, and bromine than iodine. Accordingly, we find that when chlorine acts upon the bromides they are decomposed, the bromine being expelled, and that bromine, in turn, expels iodine from combination. Therefore, such reactions as

$$KBr + Cl = KCl + Br$$
$$KI + Br = KBr + I$$
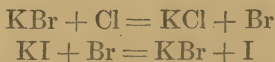
are simply illustrations of the general electrical relations of elements concerned. If these affinities were the only active causes of chemical change, the subject would be quite simple, but by repeated experiment chemists now know that the surrounding circumstances may suspend or modify the play of affinities, so as to produce an endless variety of chemical action. It is hardly necessary to remark that all the modi-

fying influences are not yet known, but a few of the more important will be considered.

Chemical action is greatly aided by reducing the substances to powder or to the liquid form. By such means we expose more surface for action, and we also secure more intimate mixture of the reactive bodies. Iron in the form of wire can be burned only with difficulty; in the form of filings it burns with some ease; in the very fine condition known as reduced iron (*q. v.*) it can be lighted with a match; and by heating the oxalate it can be obtained in so fine a form as to take fire on exposure to the air. This progressive increase of affinity is due to progressive fineness of division.

A mixture of sodium carbonate and tartaric acid will not be changed by very intimate mixture in fine powder, but as soon as thrown into water the copious escape of gas will show that the solution has brought the atoms within sufficient range. A large number of operations in assaying and metal-working are dependent upon the effects of melting in causing chemical action.

The physical forces, light, heat, electricity, etc., are frequent causes of both combination and decomposition. The respiration of plants, the fading of colors, the methods of photography, are familiar instances of the action of light. Heat is one of the most frequent agents, and electricity has been pointed out in a previous section as not only a cause of decomposition, but also as a means of determining the relation of the different elements. Besides the direct conditions which modify chemical action and disturb the simpler relations of affinity, a few influences exist which are dependent partly upon incidental circumstances and partly upon the nature of the compounds that are to be formed. These will be briefly enumerated:

(*a*) **Nascent State**. Nascent means "born," "brought forth," and expresses, when applied to an element, the fact

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that it has just been set free. In such a condition its affinity may be very much higher than ordinarily. Arsenic, for instance, does not combine with hydrogen when the two are brought together from different vessels, but if we arrange the apparatus so that the hydrogen is set free—that is, becomes "nascent"—in presence of the arsenic, a compound, AsH_3, is formed rapidly. Very extended use is made of this influence of the nascent state in forming compounds.

(*b*) **Insolubility.** When in any liquid we bring together substances which are capable of forming a body insoluble in the liquid, that insoluble compound will be produced in spite of the general relations of affinities. This influence of insolubility is the basis of a large number of tests and other chemical operations.

When the formation of the insoluble compound would require a powerful chemical agent to be set free, the change will not take place unless, of course, the added substance is stronger than the one to be liberated. Carbonic acid forms with calcium a body quite insoluble in water, but this body cannot be formed by passing carbonic acid into calcium sulphate. The reason is shown at once on examining the conditions of the experiment. The reaction would have to be $CaSO_4 + H_2CO_3 = CaCO_3$ (insoluble) $+ H_2SO_4$; that is, sulphuric acid would be set free. The affinity of H_2CO_3 is, under ordinary conditions, so much below that of H_2SO_4 that the former will not drive out the latter. The condition becomes changed if we assist the action of the carbonic acid by some substance which has an affinity for sulphuric acid and will prevent its being set free. $CaSO_4 + Na_2CO_3$ will produce immediate action, resulting in $CaCO_3 + Na_2SO_4$. This reaction illustrates a common method of keeping the powerful affinities in abeyance, and thus allowing secondary influences full play. Some of the arsenic tests show the principle strikingly. Arsenous acid added to $CuSO_4$ produces no action, because the affinity of the SO_4 is too strong,

but by adding a little ammonia the strong affinity this has for SO₄ assists in breaking up the copper sulphate, and immediately a precipitate of copper arsenite falls.

(*c*) **Volatility.** This is the second influence that disturbs ordinary affinities. If a body is capable of being converted into a gas, this fact will diminish its chemical power; fixed substances that have ordinarily less affinity will drive it out of combination. Boric acid, for instance, is one of the weak acids, yet at a red heat it will drive out even sulphuric acid. The cause is, in the main, that at this temperature sulphuric acid is volatile, while boric acid is fixed. Chemists make, as is well known, much use of the action of heat as a modifier of chemical action, and frequently it is this influence of volatility which is brought into play.

The ease with which hydrogen is driven out of combination may be regarded as due to its volatility, it being a gas even at low temperatures.

(*d*) **Mass.** Sometimes chemical action seems to be governed by the quantity of the substance present. If we pass water vapor over red-hot iron, iron oxide is formed and hydrogen is set free; if we pass the hydrogen back over the iron oxide, steam is formed and iron set free. In the first case the water is in excess and exerts an oxidizing influence; in the second, the hydrogen is in excess and exerts a deoxidizing influence. The effect of mass is indefinite and uncertain, and need not enter into the ordinary working of reactions.

It will seen to be a deduction from these statements that no substance can be set down as absolutely the strongest in affinity. It cannot be determined what is the strongest acid or the strongest alkali, except under specified conditions.

2d. The proportion in which bodies react is usually strictly according to their atomicities. Let it be required to write the reaction between mercuric chloride and potassium iodide. The formulæ are HgCl₂ and KI, but the bodies will not react
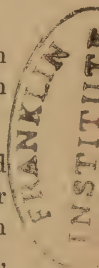
in this proportion, for the Hg will require I_2, and Cl_2 will require K_2. The proper reaction is $HgCl_2 + 2KI = HgI_2 + 2KCl$. In the same way, antimony sulphide and hydrochloric acid can only act upon each other in the ratio $Sb_2S_3 + 6HCl$, because, Sb being a triad, Sb_2 will combine with Cl_6, and S being a dyad, S_3 will require H_6.

3d. If a chemical change occurs when two given substances are brought in contact, its nature will depend principally upon the electrical relations of the elements concerned. In the reaction $HgCl_2 + H_2S$ the only possible result is the combination of S with Hg and H with Cl, as is shown at once by placing the proper signs over the elements, $\overset{+}{Hg}\overset{-}{Cl_2} + \overset{+}{H_2}\overset{-}{S}$. Such a combination as $\overset{+}{Hg}\overset{+}{H_2}$ or $\overset{-}{Cl_2}\overset{-}{S}$ could not take place, since it requires like electricities to attract, which is against the rule. In beginning with reactions, the student will do well to place the proper signs over each element, and these signs will be a useful guide and control. When substances containing three elements are part of the reaction, the signs may be placed thus:

$$\overset{+}{Ba}(\overset{-}{NO_3})_2 + \overset{+}{K_2}\overset{-}{SO_4} = \overset{+}{Ba}\overset{-}{SO_4} + 2\overset{+}{K}\overset{-}{NO_3}.$$

The placing of the single sign over the two elements is simply an evidence of the fact that in ordinary reactions these act as a single element.

The following formulæ will further illustrate the general principle:

$$\overset{+}{Ag}\overset{-}{NO_3} + \overset{+}{Na}\overset{-}{Cl} = \overset{+}{Ag}\overset{-}{Cl} + \overset{+}{Na}\overset{-}{NO_3}$$

$$\overset{+}{H_2}\overset{-}{O} + \overset{+}{K_2} = \overset{+}{K_2}\overset{-}{O} + \overset{+}{H_2}$$

$$\overset{+}{H_2}\overset{-}{O} + \overset{-}{Cl_2} = \overset{+}{H_2}\overset{-}{Cl_2} + \overset{-}{O}$$

$$\overset{+}{Ba}\overset{-}{Cl_2} + \overset{+}{K}\overset{+}{H}\overset{-}{SO_4} = \overset{+}{Ba}\overset{-}{SO_4} + \overset{+}{K}\overset{-}{Cl} + \overset{+}{H}\overset{-}{Cl}$$

In the last reaction the electro-positives K and H may seem to be in union, but this is not the case. Each is independently united to the SO_4, which is a dyad. The formula might be written $\overset{+}{\underset{H}{K}}\overset{-}{} SO_4$

In writing reactions in which any element in the free state is expressed, it is now customary to use such proportions as will give an even number of atoms of the element. Thus, in giving the reaction of sodium upon hydrochloric acid, it would be written

$$Na_2 + 2HCl = 2NaCl + H_2,$$

and not

$$Na + HCl = NaCl + H.$$

This system has been adopted in deference to the theory that no atom exists alone, but that even elementary bodies have their atoms united in pairs. No advantage of any practical character accrues to the student from this complication; and if the reaction is known and understood in the simpler form, an easy multiplication will change it into the other.

8. ACIDS, BASES AND SALTS.

THESE terms, introduced in the earlier days of chemical science, when the composition of bodies was but imperfectly understood, are still retained, but with very vague and uncertain meaning. The old definition of an acid was a body having a sour taste, a power to affect vegetable colors, especially to turn blue colors red, and forming definite compounds called *salts*. Lavoisier had advanced the doctrine that all acids contained oxygen, but the existence of sulphur acids and the acids of the chlorine group upsets this view. The study of the anhydrides, or, as they were first called, anhydrous acids, showed that water was necessary to the acid

4 *

condition, and thus was brought about the present view that hydrogen is the essential element of an acid. The explanation of these relations is given elsewhere.

Bases were defined to be oxygen compounds, capable of uniting with an acid and neutralizing it. This definition was subsequently modified by including sulphur, selenium and tellurium as capable of forming bases.

Salts were defined as bodies formed by the action of an acid upon a base. The reactions

$$CaO + H_2SO_4 = CaSO_4 + H_2O,$$
$$CaO + 2HCl = CaCl_2 + H_2O,$$

are instances of such effects. The above definition, however, does not include the production of a salt by direct action of a halogen, or of an acid, upon a metal; thus:

$$Zn + H_2SO_4 = ZnSO_4 + H_2$$
$$Zn + Cl_2 = ZnCl_2.$$
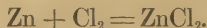
Intimately connected with this subject is the meaning of the terms ACID, ALKALINE and NEUTRAL, as applied to the conditions of substances. If we add a drop of sulphuric acid to a solution of the coloring matter of purple cabbage, the color will change to red; by the addition of a small amount of soda the color will be restored, and by further addition will be changed to green. The soda is a base; it has combined with the acid and deprived it of its chemical activity. By this combination the soda has also been neutralized, and it is only by adding it *in excess*, that we can get its specific action on the color. Other coloring principles show similar effects.

Litmus is a red color that becomes blue on the addition of a base, and of course has the red color restored on the addition of an acid. It is usually sold in the blue condition, and is used either in solution in water or in the form of *litmus-paper*—strips of paper soaked in the solution and dried.

Cochineal is much used as a substitute for litmus; acids turn it orange-yellow, and bases turn it purplish.

It would seem that by the use of these tests we could determine whether a substance was an acid, a base or a salt, but, unfortunately, the reactions just given apply only to cases in which acids and bases of similar intensity are combined. When soda, which is a strong base, is united to sulphuric acid, which is a strong acid, the compound is neutral; but the union of soda with a feeble acid like boric produces a salt which is alkaline; and the union of copper oxide with sulphuric acid gives a body which is acid to the tests mentioned above. These color reactions are of some importance in practical chemical operations, but they have no value in determining the theoretical relations between acids, bases and salts.

Salts may be divided into four classes:

NORMAL SALTS, in which the hydrogen of the acid is replaced by a single element, combined according to its atomicities. The acids themselves are normal salts:

H_2SO_4	Hydrogen sulphate (sulphuric acid).
Na_2CS_3	Sodium sulpho-carbonate.
KNO_3	Potassium nitrate.

MIXED SALTS, in which two or more elements have replaced the hydrogen. When some hydrogen remains, the body is usually called an *acid* salt:

$HKCO_3$	Acid potassium carbonate.
$KNaC_4H_4O_6$	Sodio-potassium tartrate.

It must be noticed that K_2CrO_4 is not a mixed salt; the chromium has not replaced any of the hydrogen of the acid.

DOUBLE SALTS, in which two complete salts of either of the above classes unite to form a definite compound which is generally distinctly crystalline:

$K_2SO_4 + Al_2(SO_4)_3$	Potassium and aluminum sulphate.
$BaCO_3 + CaCO_3$	Barium and calcium carbonate.
$2KCl + PtCl_4$	Potassium and platinum chloride.

CONJUGATED SALTS, in which a definite salt is united with a body not a salt. A great variety of these is known, and many are of uncertain composition. Two important classes may be recognized:

(a) *Oxy-salts*, called frequently *basic* or *sub*-salts, in which a basic oxide is united with the salt, thus:

$Bi(NO_3)_3 + Bi_2O_3$	Bismuth oxynitrate.
$SbCl_3 + Sb_2O_3$	Antimony oxychloride.

(b) *Anhydro-salts*, called frequently *acid* or *bi*-salts, in which an anhydride is united with the salt, thus:

$K_2CrO_4CrO_3$	Potassium anhydrochromate.
$2NaBO_2 + B_2O_3$	Sodium anhydroborate.

Salts formed by one atom of a monad are called MONO-BASIC; by two atoms of a monad, DIBASIC; by three atoms, TRIBASIC; by four, TETRABASIC, thus:

Metaphosphoric acid HPO_3 is monobasic.
Sulphuric acid H_2SO_4 is dibasic.
Sodium sulphantimonate Na_3SbS_4 is tribasic.
Potassium silicate K_4SiO_4 is tetrabasic.

9. QUASI-ELEMENTS, OR RADICLES.

QUASI is a Latin word meaning "as if." QUASI-ELE-MENTS are molecules which have the power of forming compounds "as if" they were elements. They are also called RADICLES (often incorrectly spelled *radicals*), because they may be looked upon as the root or basis of the compounds into which they enter. Practically, any unsaturated mole-

cule may be regarded as a radicle. Some of the important ones to which special names have been given are—

CN,   cyanogen, electro-negative monad, resembling the chlorine group.

HO,   hydroxyl, a monad combining with both negative and positive bodies.

H_4N,   ammonium, electro-positive monad, related to the potassium group.

By deducting the hydrogen from acids we get their radicles. Thus:

$$\text{The radicle of } H_2SO_4 \text{ is } SO_4.$$
$$\text{``    ``   `` } HNO_3 \text{ is } NO_3.$$

As acids are saturated molecules, the atomicity of such radicles will be equal to the number of hydrogen atoms which have been taken away.
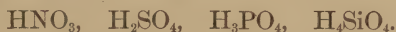
NO_3 is therefore a monad radicle,
SO_4 ``    dyad    ``
PO_4 ``    triad    ``
SiO_4 ``    tetrad    ``

because the respective acids are

$$HNO_3, \quad H_2SO_4, \quad H_3PO_4, \quad H_4SiO_4.$$

Organic chemistry presents us with a large number of radicles, the majority of which contain carbon and hydrogen.

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10. COMBINATION BY VOLUME.

In the section on the laws of chemical combination mention was made of the fact that the elements combined with each other in definite proportions, generally in the proportion of the atomic weights or some multiple of them. Modern chemical research has called attention to important facts in

regard to the proportion by *volume* or *bulk* in which the combination takes place. As yet, no very exact or important results have been obtained from the study of the volume in the solid or liquid form, but in the state of gas or vapor the relations of the different elements are quite striking.

If we weigh equal volumes of the elements in the state of gas, we find that their relative weights will, with a few exceptions, be in exact proportion to their atomic weights. For instance, a vessel which holds 1 grain of hydrogen (about 47 cubic inches) will hold the following quantities of other elements, it being understood that all the bodies are in the state of gas and at the same temperature and pressure:

Element.	Atomic Weight.	Wt. of vol., equal to 1 vol. of H.
O	16	16
S	32	32
Cl	35.4	35.4
I	127	127
Br	80	80

Some of the elements cannot be converted into vapor, and consequently cannot be compared on this system. Among these are carbon, silicon and many of the common metals. These practically resist the action of the highest temperature which can be used in such experiments. A few elements show results which are exceptional; among these are—

Element.	Atomic Weight.	Wt. of vol., equal to 1 vol. of H.
As	75	150
P	31	62
Hg	200	100

In the case of phosphorus and arsenic the weight is twice as great as what analogy would require; in the case of mercury, half as great.

Changes in temperature and pressure produce changes in the volume of gases, and all gases are affected nearly equally. It is pointed out elsewhere that a change of volume is due

to a separation of the molecules or atoms, and not to any change in the volume of the atoms themselves. It is a reasonable assumption that if two gases expand equally under the same conditions, it is because they contain the same number of interspaces in which the expansion takes place. Hence a law, which is expressed thus:

Equal volumes of elementary gases contain equal numbers of molecules.

The weight of the atoms of each element may be determined by this law. If a given volume of hydrogen contains, say, 1000 molecules, the same volume of oxygen will contain the same number; and as the oxygen volume is 16 times as heavy as the hydrogen, it is clear that the weight of each molecule of oxygen will be 16 times that of each molecule of hydrogen. The molecules of hydrogen and oxygen each contain two atoms, hence the atomic weights will also be in the proportion of 16 to 1.

As gases are decidedly affected by even slight changes in temperature and pressure, it becomes necessary to refer all observations to a standard condition. Under the English system of weights and measures this has usually been 60° F. for temperature, and 30 inches of mercury (about 15 pounds per square inch) for pressure. Under the French system the temperature is 0° C. (sometimes 15° C.) and 760 millimetres of mercury. This latter figure is practically the same as 30 inches, and is the height of the barometer at the level of the sea.

A study of the condition of the ordinary gases leads us to believe that the spaces between the molecules are very much greater than the molecules themselves. Ample room exists in any gas for adding atoms or molecules without increasing the volume. The phenomena of combination between gases seem to show that all molecules are of the same size; at least it is known that the elementary gases combine in such a way as to produce a volume of gas *which is equal to twice the vol-*

ume that would be occupied by one atomic weight of hydrogen.
The following instances are taken from among the commonest
chemical compounds:

One volume of H and one volume of Cl combine and pro-
duce *two* volumes of HCl.

Two volumes of H and one volume of O combine and pro-
duce *two* volumes of H_2O.

Three volumes of H and one volume of N combine and
produce *two* volumes of NH_3.

In each of these cases it is seen that the volume of the
resulting combination is twice that of the one atomic weight
of hydrogen. If the weights should be in grammes, then the
resulting volume would be that occupied by two grammes of
H; if the weights should be in grains, then the resulting vol-
ume would be that occupied by two grains of H; and if the
substances were estimated by volume alone, say in pints, then
the resulting compounds would have the volume of two pints.

The great majority of chemical compounds that have been
examined conform to this law. A few exceptions are known,
and, as mentioned above, some elements cannot be converted
into gas, and thus cannot be examined on this point. It fol-
lows, from this uniform condensation, that when the molecule
contains many atoms the gas will be heavy, and, further, that
its weight compared to hydrogen can be easily calculated.

Some examples will make this plain:

47 cubic inches of H, weighing 1 grain, will combine with
47 cubic inches of Cl, weighing 35.4 grains, and produce 94
cubic inches (*i. e.* 47 × 2) of hydrochloric acid (HCl), weigh-
ing 36.4 grains; and by dividing this last result by 2 we get
the weight of a quantity of hydrochloric acid equal to the
one atomic weight of hydrogen—viz. 18.2. This figure, 18.2,
represents, therefore, the density or specific gravity compared
to hydrogen.

94 cubic inches of H, weighing 2 grains, will combine with
47 cubic inches of O, weighing 16 grains, and produce 94

cubic inches of steam, H_2O, weighing 18 grains. If we divide 18 by 2, we get, as before, the density of steam compared to hydrogen—viz. 9.

47 cubic inches of N, weighing 14 grains, will combine with 141 cubic inches (47×3) of H, weighing 3 grains, and form 94 cubic inches of ammonia, NH_3, weighing 17 grains; and this weight, divided by 2, gives 8.5 as the density of ammonia compared to hydrogen.

These principles are employed in determining the formulæ of bodies. N and O combine to form a body called nitric oxide, which is sometimes written NO and sometimes N_2O_2. The following calculation will show which is correct:

The formula NO requires

One volume of N $= 14$
    "   "   " O $= 16$
              $\overline{30}$ $30 \div 2 = 15$, producing two volumes NO.

The formula N_2O_2 will require

Two volumes of N $= 28$
    "   "   " O $= 32$
              $\overline{60}$ $60 \div 2 = 30$, producing two volumes N_2O_2.

In the first instance the formula would indicate a vapor fifteen times as heavy as hydrogen; in the second case, thirty times as heavy. Experiment shows the first number to be correct, and therefore justifies the formula NO.

11. RELATION BETWEEN SPECIFIC HEAT AND ATOMIC WEIGHT.

THE SPECIFIC HEAT of any substance is the proportion between the amount of heat required to raise the body to

any temperature, and that required to raise the same weight of water to the same temperature. One pound of water requires thirty-three times as much heat to raise it to any temperature as mercury does; the specific heat of mercury is therefore $\frac{1}{33}$.

Many of the elements show the interesting fact that multiplying their specific heat and atomic weight together gives nearly similar numbers, as the following table shows:

Element.	Atomic weight.	Specific heat.	Product.
Na	23	0.2934	6.75
S	32	0.2026	6.48
Cu	63.4	0.0952	6.04
Zn	65.2	0.0956	6.24
As	75	0.0814	6.10
Br	80	0.0843	6.75
Au	197	0.0324	6.38
Hg	200	0.0319	6.38
Pb	207	0.0314	6.50

Some of the elements give different results in different conditions, and some give results that are entirely at variance with the law.

12. NUMERICAL RELATIONS OF THE ATOMIC WEIGHTS.

INDEPENDENTLY of the relations between the atomic weights and other quantities discussed in the preceding sections, several attempts have been made to exhibit the relations between the atomic weights themselves, especially between the weights of bodies belonging to the same group. Some of these attempts have been passed over as merely curious speculations; others have been regarded as of value in classifying the elements, and even in deciding between differing determinations of atomic weight. Among the most recent and most noticeable of these systems is the one known as

Mendelejeff's Law of Periodicity.—It has recently attracted considerable attention, because it is supposed to point out the true classification of the elements, and to give indication of elements existing but yet undiscovered. The arrangement, as presented by Mendelejeff, is best shown by the table on the next page.

Many gaps, however, are noticed ; some elements are placed out of the commonly-accepted arrangement ; others have new atomic weights assigned to them. The element gallium was not known at the time this system was published, and the fact that it fits one of the gaps is regarded as a proof of the value of the tables. Mendelejeff has expressed the points of his system in a mathematical phrase, as follows :

The properties of elements, the constitution and properties of their combinations, are periodic functions of the atomic weights.

Series	First Group.	Second Group.	Third Group.	Fourth Group.	Fifth Group.	Sixth Group.	Seventh Group.	Eighth Group.
1	H = 1							
2	Li = 7	Be = 9	B = 11	C = 12	N = 14	O = 16	F = 19	
3	Na = 23	Mg = 24	Al = 27	Si = 28	P = 31	S = 32	Cl = 35	
4	K = 39	Ca = 40	?E = 44	Ti = 48	V = 51	Cr = 52	Mu = 55	Cu = 63, Fe = 56, Co = 59, Ni = 59
5	Cu = 63	Zn = 65	Ga = 69.9	? = 72	As = 75	Se = 78	Br = 80	
6	Rb = 85	Sr = 87	Y = 88	Zr = 90	Nb = 94	Mo = 96	? = 100	Ru = 104, Rh = 104, Pd = 106, Ag = 108
7	Ag = 108	Cd = 112	In = 113	Sn = 118	Sb = 122	Te = 125	I = 127	
8	Cs = 133	Ba = 137	Di = 138(?)	Ce = 140	....	....	......	Os = 195, Ir = 197, Pt = 198, Au = 199
9	....	....	Er = 178	La = 180?	Ta = 182	W = 184	? = 190	
10								
11	Au = 199	Hg = 200	Tl = 204	Pb = 207	Bi = 208	U = 250		
12				Th = 231	....			

DESCRIPTIVE CHEMISTRY.

A COMPLETE table of the elements, their atomicities, atomic weights and symbols, will be found at the end of the book. Before beginning Descriptive Chemistry a brief account will be given of the important groups.

1. The **Oxygen Group** includes oxygen, sulphur, selenium and tellurium. They are electro-negative dyads, and possess the power of forming, with many elements, basic or acid compounds, according to the proportion in which they are combined.

2. The **Chlorine Group** includes fluorine, chlorine, bromine and iodine. They are electro-negative monads, and are the only elements which form salts without the aid of some member of the oxygen group. For this reason they have been called the HALOGENS, a word meaning "salt-formers."

3. The **Nitrogen Group** includes boron, nitrogen, phosphorus, arsenic, antimony, and probably vanadium, bismuth and gold. They are of uneven atomicity, generally triads or pentads; their electrical relations are intermediate in character, neither strongly positive nor strongly negative. They form anhydrides, distinguished by the power of combining with water in two or more proportions, forming distinct acids. Several of them combine easily with hydrogen to produce bodies which have analogies to the members of the potassium group (*q. v.*).

4. The **Carbon Group** includes carbon, silicon, titanium,

tin and some rarer elements. They are tetrads, and, like the nitrogen group, their electrical relations are intermediate. Boron is sometimes classed here, but it is a triad and belongs to the nitrogen group.

5. The **Potassium Group** includes hydrogen, lithium, sodium, potassium, rubidium, cæsium and silver. They are electro-positive monads, and—with the exception of hydrogen and silver, which differ in several material points from the rest of the group—have high affinity for the members of the oxygen and chlorine groups. With oxygen they produce powerful corrosive bases called the alkalies, and on this account are sometimes called the alkali metals. Hydrogen and silver are the only ones that occur free in nature.

6. The **Calcium Group** includes calcium, barium, strontium and lead. They are electro-positive dyads, and form oxides which are slightly soluble in water, but much less caustic or corrosive than the alkalies proper, and are often called alkaline earths. Their sulphates, carbonates and phosphates are practically insoluble.

7. The **Zinc Group** includes zinc, magnesium, cadmium and beryllium. They are never found free, but are tolerably easily reduced from their compounds. They are electro-positive dyads, and form a definite oxide which is insoluble in water, not caustic, but capable of forming well-marked salts.

8. The **Iron Group** is electro-positive, and includes aluminum, iron, manganese, nickel, cobalt, chromium and probably several other elements the chemistry of which is not well known. They are not found in the metallic state, except in small quantity. Most of them form two sets of compounds, acting in one as dyads, in the other as double tetrads. Several form well-marked anhydrides.

9. The **Copper Group** includes copper and mercury, electro-positive dyads, resembling each other in the power of forming

two sets of compounds, in one of which they act as dyads, and in the other *apparently* as monads. In the apparently monad condition they form chlorides insoluble in water, and are in this way partly related to silver.

10. The **Platinum Group.** A number of elements which are found in association with platinum are usually grouped together under the name of platinum metals. These are platinum, palladium, iridium, rhodium, ruthenium and osmium. They do not, however, appear to be connected by any striking relations, and when their properties and compounds become better understood they will probably be distributed among the groups already described. Pt, Pd and Ru agree in some respects with tin ; Ir and Rh with iron ; Os with nitrogen.

Unclassified Elements. Some of the elements are either so rare that their relations have not yet been satisfactorily studied, or their properties are such as to render it impossible to classify them satisfactorily under any system. Among these are molybdenum, zirconium, tantalum, tungsten, yttrium, thorium, erbium, uranium, cerium, lanthanum, didymium, gallium, indium and thallium. The brief descriptions that these need will be given in connection with those groups to which they seem to be related.

An arrangement in series like that just given can never show the true relations of the elements. By a tabular form, such as is given on page 56, each element can be placed so as to show its relations to the others. In the table the elements that are distinctly electro-negative are printed in bold type; those distinctly electro-positive, in ordinary Roman type ; those intermediate, in *italic.*

CLASSIFICATION OF THE ELEMENTS.

Electro-negative in **bold type**; intermediate, in *italic*; electro-positive, in Roman.

H, Li, K, Na, K, Rb, Cs ——— Ag

Cu
Hg

Ca, Ba, Sr, Pb ———

Th, Y, E.

Zn, Cd, Be, Mg, ———

In

Di, La
Ce, Ga

Al ——— Cr

Co
Ni   Fe   Mn
      U

Rh
Ir

Mo ——— *V, Bi, Sb, As, P, N, B* ——— C, Si, Ti, Sn, Ta, W, Zr

Os

Tl ——— Au

Pt
Pd
Ru

F
Cl
Br
I

O
S
Se
Te

Chlorine Group. This includes chlorine, bromine, iodine and fluorine. They are negative monads, and of high chemical affinity. They combine with oxygen in several proportions, assuming atomicities of three, five and seven, but the compounds are easily decomposed. Chlorine generally expels bromine from combination, and bromine expels iodine.

CHLORINE, Cl, 35.46.

Chlorine was discovered by Scheele in 1774.

The name, derived from a Greek word meaning " green," was given by Davy.

Sources. Chlorine is always found in combination. Its most abundant compound is common salt, NaCl, which is found in the animal, vegetable and mineral kingdoms. The chlorides of lead, silver and some other bodies are found as minerals. Of late years a considerable amount of potassium chloride has been obtained from the salt-mines of Stassfurth, Germany. Hydrogen chloride is occasionally found in volcanic regions.

Preparation. Chlorine being extensively used in manufacturing chemistry, many processes for its preparation have been devised; nearly all of them depend upon the oxidation of some chloride. Hydrogen chloride, hydrochloric acid, is generally used. The following include the processes best suited for the laboratory:

(*a*) By heating a mixture of manganese dioxide and hydrochloric acid,

$$MnO_2 + 4HCl = MnCl_2 + 2H_2O + Cl_2.$$
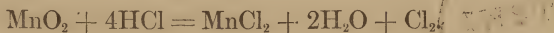
(*b*) By heating a mixture of common salt, sulphuric acid and manganese dioxide,

$$MnO_2 + 2NaCl + 2H_2SO_4 = MnSO_4 + Na_2SO_4 + 2H_2O + Cl_2.$$

(*c*) By the action of hydrochloric acid upon potassium chlorate, potassium bichromate or bleaching-powder. These

methods are suitable for the preparation of small amounts of chlorine for use as a test.

The process (b) is the most economical, (a) the most convenient for laboratory-work.
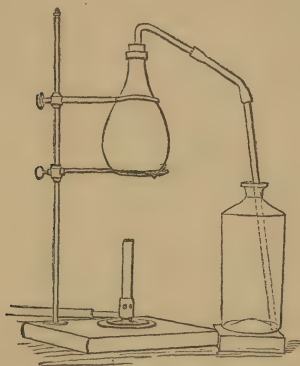
Exp. Mix in a flask about 1 ounce of black oxide of manganese with 4 ounces of hydrochloric acid; shake well, and heat the mixture gently. The chlorine is best collected by displacement. The operation should be conducted on the small scale and in a well-ventilated place, as the gas is very irritating. Narrow-mouthed, stoppered bottles of about one pint capacity will answer very well for receiving the gas. It cannot be collected over water or mercury.

Properties. Chlorine is a greenish-yellow gas of a disagreeable and highly irritating odor. By a pressure of about 60 pounds to the inch it condenses to a greenish liquid which has never been frozen. The gas is about two and a half times as heavy as air; one litre weighs 3.1808 grms.; water dissolves about three volumes, acquiring the color and odor of the gas; the solution, known as *chlorine water*, does not keep well.

The affinities of chlorine are very great. It combines with every element. It combines directly with most of the metals, decomposes water, and bleaches and destroys many organic substances. Its affinity for hydrogen is increased by light.

Exp. A lighted taper put into a jar of chlorine continues to burn, but with a dark red flame and the escape of clouds of smoke. The chlorine has combined with the hydrogen of the wax, and not with carbon, hence the abundant liberation of the latter.

Exp. Powdered antimony dropped into chlorine takes fire at once,

and produces dense, irritating clouds of antimony chloride. Dutch leaf, a sort of brass, is also instantly burned up.

Exp. Paper dipped in oil of turpentine takes fire spontaneously in chlorine, producing a red flame and a dense cloud of smoke. This experiment requires pretty pure chlorine and good turpentine.

Exp. The bleaching power of chlorine is easily shown by pouring some solution of cochineal or litmus into a jar of the gas. The color is almost instantly removed. Pieces of calico may also be quickly bleached by placing them, in a wet state, in contact with the gas.

Exp. A small quantity of the gas as it comes from the evolution-flask should be allowed to bubble through water. If a small quantity of this liquid be shaken in a bottle containing hydrogen sulphide, the odor of this gas will disappear.

The last two experiments show the bleaching and disinfecting applications of chlorine. It is, however, rarely used in the form of gas, on account of the obvious inconvenience. It is usually employed in the form of bleaching-powder, made by passing the gas into slaked lime. The body so produced is described among the calcium salts. It is easily decomposed by dilute acids, yielding its chlorine; even the carbonic acid of the air will act upon it. It is generally employed in solution, and is often incorrectly called chloride of lime.

Chlorine does not bleach unless moist, and it is believed that a decomposition of water first occurs, and that the oxygen thus set free is the active agent,

$$H_2O + Cl_2 = 2HCl + O.$$

The powerful action of chlorine is utilized occasionally for the decomposition of ores.

Gen. Chem. Rel. In the chlorides the chlorine is electro-negative and monatomic. It is capable, however, of assuming electro-positive relations and higher atomicities, and in its compounds with oxygen shows both these changes. It is also capable of replacing hydrogen, atom for atom, giving rise to an important and extensive series of *substitution* compounds, which are considered in connection with Organic Chemistry.

The chlorides are generally less numerous than the corresponding oxides; for instance, we know two oxides of hydrogen, but only one chloride.

Hydrochloric Acid, Muriatic Acid, Spirit of Salt, HCl. This substance was long known in an impure form. Basil Valentine (fifteenth century) described the pure acid, and Davy in 1810 showed its composition.

**Sources.** Hydrochloric acid occurs in the gases evolved from volcanoes and in solution in the waters of some mountain-streams of South America.

Preparation. The acid may be formed by the direct union of its elements, but this method has only theoretical interest. The practical process is the action of common salt and sulphuric acid, according to the following reaction:

$$2NaCl + H_2SO_4 = Na_2SO_4 + 2HCl.$$

This reaction requires a high temperature. In ordinary experiments and on the small scale the reaction is

$$NaCl + H_2SO_4 = NaHSO_4 + HCl;$$

by which, from a given amount of sulphuric acid, only half the quantity of hydrochloric acid is obtained. This form of the process is therefore costly.
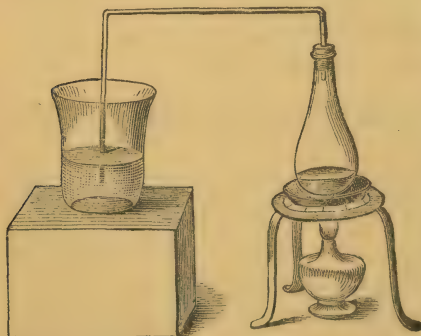
Exp. One ounce of common table salt is mixed with twice its weight of strong sulphuric acid in a flask provided with a delivery-tube in the ordinary manner as shown. The hydrochloric acid comes off freely, and may be collected by downward displacement, as in the preparation of chlorine, or passed into water. Heat may be used, but the tube must be taken from the water when the heat is withdrawn.

Properties. Hydrochloric acid is a colorless gas of a strong pungent odor and poisonous to animals and plants. Its density is 18.181; 1 litre weighs 1.63 grms. It may be liquefied by a pressure of about 600 pounds to the inch. It does not burn nor support ordinary combustion, but some substances burn in it, forming chlorides. Its most important property is its solubility in water, which at low temperature will absorb nearly 500 volumes, producing a strongly acid solution, which is the common hydrochloric or muriatic acid. The strongest form usually sold contains 43 per cent. by weight of the gas. When pure it is a colorless, fuming, strongly acid liquid, but the commercial forms are usually yellow, from the presence of iron.

Gen. Chem. Rel. Hydrochloric acid is rather weaker than sulphuric or nitric acid, but is used largely as a solvent. It generally acts by forming chlorides. When the body dissolved is an element, free hydrogen generally escapes; but when compounds are dissolved, the hydrogen usually combines with the substance displaced by the chlorine.

In this way oxides dissolve in hydrochloric acid and form chlorides and water; sulphides form hydrogen sulphide, etc. The following reactions are instances:

$$Zn + 2HCl = ZnCl_2 + H_2.$$
$$ZnO + 2HCl = ZnCl_2 + H_2O.$$
$$FeS + 2HCl = FeCl_2 + H_2S.$$

With oxides more rich in oxygen (MnO_2, Fe_2O_3, CrO_3) the action is dependent upon the temperature and other conditions. Sometimes the whole of the chlorine is retained, in other cases a portion escapes.

$$Fe_2O_3 + 6HCl = Fe_2Cl_6 + 3H_2O.$$
$$MnO_2 + 4HCl = MnCl_2 + 2H_2O + Cl_2.$$

Tests. Hydrochloric acid produces, with ammonia, white fumes of NH_4Cl. In common with all the chlorides, it pro-

6

duces, with silver nitrate, a white curdy precipitate of silver chloride, soluble in ammonia.

A mixture of about three parts nitric with five parts hydrochloric acid has been long used under the names *aqua regia* and nitro-muriatic acid. It dissolves gold and platinum, and owes its efficacy in part to the free chlorine which is formed by the oxidizing action of the nitric acid upon the hydrogen of the muriatic.

Compounds of Chlorine and Oxygen. Three of these are known in the free state:

Cl_2O Hypochlorous anhydride.
Cl_2O_3 Chlorous anhydride.
ClO_2 Chlorine dioxide.

Several others are known only in combination, and in these we have a well-marked and important series:

$HClO$ Hypochlorous acid.
$HClO_2$ Chlorous   "
$HClO_3$ Chloric   "
$HClO_4$ Perchloric   "

Hypochlorous anhydride, Cl_2O, is formed when chlorine is passed over mercuric oxide,

$$HgO + Cl_4 = Cl_2O + HgCl_2.$$
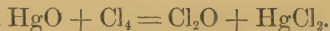
The combination of oxygen and chlorine will not take place when the two elements are mixed under ordinary conditions. The action in this case may be regarded as an instance of the influence of the nascent state. Hypochlorous anhydride is absorbed by water, and is supposed to form hypochlorous acid, $HClO$, which has been used as a bleaching agent, especially for removing ink-stains.

Exp. Shake a few grains of finely-powdered mercuric oxide with some chlorine water. The odor of chlorine will be replaced by that of hypochlorous acid, and the liquid will easily remove stains of writing without seriously injuring the paper. Other hypochlorites may be

formed by the action of chlorine upon metallic oxides or hydrates at a *low* temperature. (See *Calcium Hypochlorite.*)

Chloric Acid, $HClO_3$. If the action of chlorine upon metallic oxides or hydrates be at a temperature of over 60° F. (15.5° C.), *chlorates* will be produced, according to the following reaction:

$$6KHO + Cl_6 = KClO_3 + 5KCl + 3H_2O.$$

The chlorate and chloride are separated by difference of solubility in water. Chloric acid may be obtained from the chlorates by stronger acids, but it has no practical value.

The chlorates are useful for the large amount of oxygen which they contain, and which they give up easily when heated. Potassium chlorate is the substance from which oxygen is usually prepared. It is used largely in fireworks. Perchloric acid is obtained by heating dilute chloric acid. The perchlorates resemble the chlorates.

The other compounds of chlorine and oxygen have no practical importance.

Chlorine combines with nitrogen to form a body called nitrogen chloride, of which the composition is somewhat uncertain. It is an oily liquid, which decomposes very easily and with a violent explosion.

BROMINE, Br, 80.

Sources. Bromine occurs in sea-water and sea-plants, in brine-springs and in a few minerals. Its compounds are generally associated with those of chlorine. Considerable quantities have of late been obtained from brine-springs in western Pennsylvania. Bromine was discovered by Balard in 1826. The name is derived from a Greek word meaning "bad smell."

Preparation. Bromine is prepared by processes analogous to those of chlorine—acting upon bromides by means of

oxidizing agents, such as a mixture of sulphuric acid and manganese dioxide. It may also be directly expelled by the superior affinity of chlorine.

Exp. Dissolve some potassium bromide in a little water; add a pinch of manganese dioxide and a few drops of sulphuric acid. Heat gently, and dark red vapors of bromine will soon be evolved. The reaction is precisely like that for chlorine:

$$2KBr + MnO_2 + 2H_2SO_4 = K_2SO_4 + MnSO_4 + 2H_2O + Br_2.$$

Exp. To a solution of potassium bromide in water add a few drops of chlorine water, and shake. The bromine will be set free and color the water yellowish. If some ether be now poured in and shaken for a few moments, the bromine will leave the water and dissolve in the ether, which will form a red layer on the surface of the water. The reaction is

$$KBr + Cl = KCl + Br.$$

The free bromine is usually dissolved in potassium hydrate, by which a mixture of potassium bromate and bromide is produced. The reaction is similar to that which occurs in the preparation of potassium chlorate (page 63). This mixture being heated, the bromate loses oxygen and forms bromide.

Properties. Bromine is a dark red liquid, which at ordinary temperatures evolves red vapors of an excessively irritating and disagreeable odor. The liquid is three times as heavy as water, and boils at 145° F. (63° C.) and freezes at —12° F. (—25° C.). It is soluble in water, and is often conveniently used in that form. Its chemical properties are similar to those of chlorine, but are not so energetic. It bleaches vegetable colors, and by decomposing water acts as an oxidizing agent. A number of its compounds are used in medicine.

Exp. The affinity of bromine may be easily shown by placing a few drops on a small piece of phosphorus, which will at once be ignited.

Gen. Chem. Rel. The chemical relations of bromine are almost exactly those of chlorine. It combines energetically, forming bromides, of which those of hydrogen, potassium and

ammonium are the most important. In these compounds the bromine is a monad. It also forms oxygen compounds analogous to those of chlorine, but they have little interest.

Tests. Bromides may be recognized by the production of the red vapor of bromine by the action of free chlorine or of a mixture of manganese dioxide and sulphuric acid. With solutions of bromides silver nitrate gives a light yellow precipitate of silver bromide which is slightly soluble in ammonia.

Hydrogen Bromide, Hydrobromic Acid, HBr. This substance cannot be conveniently prepared by the action of sulphuric acid upon a bromide, which would be the theoretical method, because the sulphuric acid is decomposed by the hydrogen, and instead of the reaction

$$2KBr + H_2SO_4 = K_2SO_4 + 2HBr,$$

we have the HBr acting on another portion of the sulphuric acid and giving

$$2HBr + H_2SO_4 = 2H_2O + SO_2 + Br_2.$$
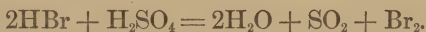
Hydrobromic acid is obtained by using a mixture of phosphorus, powdered glass and bromine, or by the action of phosphoric acid upon a bromide. It resembles hydrochloric acid in its properties, and is used in medicine.

Bromic Acid, $HBrO_3$, and **Hypobromous Acid,** HBrO, are also known. They closely resemble the corresponding chlorine compounds.

IODINE, I, 127.

Sources. Iodine occurs in association with bromine and chlorine in sea-water and sea-plants. The latter are burned, and the ashes, called *kelp*, contain various iodides.

Iodine was discovered by Courtois in 1811. The name means "violet-colored," and refers to the vapor.

Preparation. Iodine is prepared by processes similar to those for bromine, either by the action of chlorine or of a mixture of manganese dioxide and sulphuric acid. The reactions are

$$KI + Cl = KCl + I,$$

or

$$2KI + MnO_2 + 2H_2SO_4 = K_2SO_4 + MnSO_4 + 2H_2O + I_2.$$

The reactions may be shown by substituting an iodide for a bromide in the experiments described on page 64.

Properties. Iodine forms bluish-black crystalline masses with a metallic lustre. It evaporates slowly at ordinary temperatures, melts at 225° F. (107° C.), and boils at 347° F. (175° C.). The vapor has a deep violet color and a peculiar odor, somewhat like that of chlorine, but not so irritating. The solid dissolves slightly in water—much more freely in water containing potassium iodide, and in alcohol, ether and carbon disulphide. It has some bleaching and oxidizing powers. One of its important properties is the power of producing a blue color with starch. For this action the iodine must be in the free state; the iodides give no color.

Exp. Prepare some starch solution by boiling common starch with enough water to make a thin liquid. To one portion of this liquid add a few grains of potassium iodide. To another portion add a drop or two of solution of iodine. The first solution will remain colorless; the second will become deep blue. Dip slips of paper into the colorless solution, and expose them to either the vapors of nitric acid, chlorine or ozone; a blue color will be immediately produced, because the iodine is set free. Papers prepared in this manner are therefore used as tests for the vapors just mentioned.

Exp. The affinity of iodine is shown by adding a few grains to a small piece of phosphorus. Combustion occurs in a few seconds, and if a large bell jar be set down over the mass a quantity of iodine vapor will be collected in it.

Gen. Chem. Rel. The chemical relations of iodine are substantially the same as those of chlorine and bromine.

Tests. Iodides are recognized by the production of a violet vapor when treated with a free chlorine or a mixture of manganese dioxide and sulphuric acid. A solution of silver nitrate gives, with soluble iodides, a yellow precipitate, silver iodide, insoluble in ammonia. Free iodine is easily recognized by its action on starch.

Hydriodic Acid, Hydrogen Iodide, HI. This is prepared by methods similar to those used for hydrogen bromide, which body it closely resembles. It is used in medicine.

Two compounds with oxygen are known, of which iodic acid, HIO_3, has some little importance from being used as a test for morphia, which produces with it a brown color.

By the action of strong ammonia upon powdered iodine a brownish substance is produced, which was supposed to be nitrogen iodide, but probably contains hydrogen. It is easily handled while wet, but when perfectly dry explodes, with a loud report, on the slightest touch.

Exp. Put a few grains of iodine into a watch-glass and pour on enough strong ammonia to cover the mass. Crush the iodine with a glass rod, and allow the mixture to remain quiet for five minutes; then pour off the ammonia, and put the brown powder in small portions on pieces of filter-paper. After about half an hour the powder will be dry, and will explode on the slightest touch.

FLUORINE, F, 19.

Sources. Fluorine is tolerably abundant as fluor spar, CaF_2, and cryolite, $6NaF$, Al_2F_6, and some rarer minerals. It exists in the stems of grasses and in bones and teeth.

Preparation and Properties. Fluorine has never been satisfactorily prepared. Its high affinities and its special power of acting on glass and on metals render it difficult to experiment with it. It has been described as a yellow gas, and also as colorless.
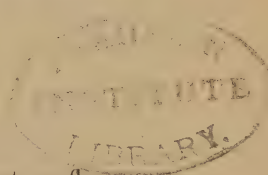
Gen. Chem. Rel. It combines with every known element except oxygen, and bears close resemblance to chlorine in most of its relations. It is remarkable for its affinity for silicon.

Hydrogen Fluoride, Hydrofluoric Acid, HF. This body is easily prepared by acting on calcium fluoride, CaF_2, with sulphuric acid. The operation must be performed in vessels of lead or platinum. The pure HF is a gas, but it is commonly seen as a strong solution in water. It acts powerfully, especially on siliceous materials. It is used for etching designs on glass.

Exp. Prepare a glass plate, flowing over it some wax or paraffine, and cutting away the coating in any design. Invert the coated plate over a leaden dish in which have just been placed some strong sulphuric acid and enough powdered calcium fluoride to make a thick paste. The action may be assisted by a gentle heat, but care must be taken not to melt the wax. After ten or fifteen minutes the uncovered parts of the glass will be found corroded. The wax can be gotten off by warming the plate.

Strong solution of hydrogen fluoride is now sold in gutta-percha bottles, upon which it has no action.

Tests. Fluorine compounds are recognized by their power of producing hydrogen fluoride when heated with sulphuric acid. The hydrogen fluoride is easily detected by its action on glass.

Potassium Group. The potassium group proper includes potassium, sodium, lithium, rubidium, cæsium. They are electro-positive monads, and are of such high affinities that they never occur in nature in the free state. Their compounds are nearly all soluble in water. Their oxides and hydrates are powerfully corrosive, and are known as the caustic alkalies. Hydrogen and silver, being positive monads, are also classed in this group, although they differ from the rest in some points.

HYDROGEN, H, 1.

Sources. Hydrogen exists in water and in all animal and vegetable substances. It was discovered by Cavendish in 1776. The name means "producer of water."

Preparation. 1. By the action of an electrical current on water or various acids. The familiar experiment of passing a current through water acidulated with sulphuric acid, although called "the decomposition of water," is really the decomposition of the acid. Pure water requires a very powerful current.

2. By the action of elements of high affinity on water or acids. With acids the action generally occurs without the aid of heat; with water, sodium and potassium act in the cold; iron, magnesium, zinc, etc. require a high temperature. Carbon may also be used, and is especially suitable when very large quantities are required.

3. By the action of alkaline solutions on a mixture of zinc and iron filings or on aluminum.

4. By heating a mixture of potassium formate and potassium hydrate,

$$KCHO_2 + KHO = K_2CO_3 + H_2.$$
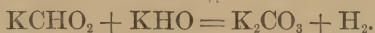
This method yields a pure gas, and was used by Pictet in his experiment on its liquefaction.

Exp. A piece of sodium amalgam (*q. v.*) is held in a spoon under an inverted test-tube filled with water and standing on the pneumatic trough. The gas collects in the tube, and may be tested as explained below. The water becomes alkaline from the formation of caustic soda,

$$Na + H_2O = NaHO + H.$$
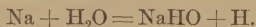
Exp. Magnesium ribbon is folded once or twice, and placed in a small hard-glass tube, one end of which is partially drawn down to a jet, the other attached to an apparatus for producing steam. Steam is passed through until the tube and ribbon are free from condensed moisture. The metal is then strongly heated at the extreme point; it takes

fire, producing an oxide and free hydrogen, which latter can be burned at the jet.

$$Mg + H_2O = MgO + H_2.$$

Exp. Fragments of zinc are introduced into the gas bottle, and dilute sulphuric or hydrochloric acid added; the gas comes off freely.

$$Zn + H_2SO_4 = ZnSO_4 + H_2.$$

Being very light and insoluble in water, it may be collected either by the pneumatic trough, as shown under oxygen, or by upward displacement; that is, allowing the tube to touch the bottom of a wide-mouthed bottle or jar inverted over the jet.

A mixture of hydrogen with air explodes when lighted, hence no experiment should be made until all the air is driven out of the evolution-bottle. The best test is to collect a small quantity in an inverted test-tube by displacement. After the gas has passed in for about half a minute, the tube should be removed, and, still keeping it inverted, a light should be applied to its mouth; a sharp explosion, extending at once through the tube, indicates the presence of air; a slight explosion at the mouth of the tube only, and a faint flame moving gradually up the tube, show the gas to be pure.

Properties. Pure hydrogen is colorless, tasteless and odorless. It is the lightest body known, a litre weighing 0.08961 grm. 100 cubic inches weigh 2.14 grains. It can be liquefied only by intense cold and pressure.

The important properties of hydrogen may be shown by the following experiments:

1. Lighted at the mouth of the delivery-tube, it burns with a pale blue, but very hot, flame. If the jet is of glass, the flame becomes yellow; a dry vessel held over it becomes coated with moisture, which is the sole product of the combustion.

2. A stout wide-mouthed bottle, or better a small tin vessel, filled with a mixture of the gas and air gives a loud explosion on being ignited.

3. Soap-bubbles blown with the gas rise rapidly in the air.

4. A large light vessel being counterpoised on a delicate balance will be thrown decidedly out of balance by substituting hydrogen for the contained air by displacement. On placing the vessel with the mouth upward, the hydrogen quickly escapes and the equilibrium is restored.

Hydrogen, though not poisonous, will not sustain life; small quantities can be inhaled without danger, and produce a peculiar change in the voice. For this the gas must be absolutely pure, which that made from commercial articles never is.

Gen. Chem. Rel. Hydrogen is electro-positive, and is a standard for atomicity, atomic and molecular weight and density. It combines with many elements. As explained in the section on nomenclature, it is regarded as the essential element of acids. Keeping in view the fact that the affinity of substances is diminished by volatility, it will be seen that the hydrogen compounds should be easily decomposable, and should be active chemical agents. Such a body as H_2SO_4 is the most active of the sulphates, because its positive element (hydrogen) is of low affinity, and is displaced by a large number of bodies.

Water, H_2O. When two volumes of hydrogen and one volume of oxygen are combined, complete condensation takes place and water is formed. When water is decomposed by the electrical current, two volumes of hydrogen and one volume of oxygen are always obtained. Since oxygen is sixteen times as heavy as hydrogen, the proportion by weight will be 2 to 16 or 1 to 8. The formula H_2O indicates that the molecule contains two atoms of hydrogen and one atom of oxygen. When the combination of the two gases is performed at a temperature of 212° F. (100° C.), the water remains as steam, equal in volume to the original hydrogen. The theoretical relations of this fact have been considered in the section on combination by volume. The composition of water *by weight* has been established with great accuracy by passing hydrogen over hot copper oxide; the oxygen is abstracted from the copper and combines with the hydrogen. The water so formed is collected and weighed; the loss of weight of the copper oxide gives the quantity of oxygen taken up, and the difference is the hydrogen. The actual result of

such an experiment, conducted with every precaution to insure accuracy, is—

Hydrogen	11.11
Oxygen	88.89
	100.00

which is substantially 1 to 8.

Properties. Pure water is odorless and tasteless, and in large masses distinctly blue. Under ordinary pressure it freezes at 32° F. (0° C.) and boils at 212° F. (100° C.). It dissolves many bodies, solid, liquid and gaseous, and is absorbed by many substances. Porous bodies, for instance, exposed to moist air increase in weight considerably by absorbing water. These effects are not generally regarded as due to chemical affinity. Solution of solids in water is generally attended with the production of cold, as may be shown by making a solution of Epsom salt or ammonium chloride. The solution of gases is apt to be attended with the production of heat. The solvent power of water is quite extensive, though very different for different bodies. As a rule, the solubility of solids is increased by heat, that of gases diminished. Gases exhibit great differences; some—*e. g.* hydrogen and oxygen—dissolve very sparingly; ammonia and hydrochloric acid, on the other hand, are largely dissolved. Gases may all be expelled by thorough boiling. When strong solutions are prepared by heat, the liquid on cooling usually deposits a portion of the dissolved substance in a crystalline form.

Exp. Bring some water to boiling, and introduce, by small portions at a time, potassium chlorate until no more dissolves; pour the solution into a thin glass beaker, and the chlorate will be continually deposited in fine scales until the liquid is quite cold. Lead iodide treated in the same way gives even a better effect.

Water is extensively distributed in nature. Besides being collected into large masses, oceans, lakes, etc., it exists in suspension in the atmosphere, in most minerals and in all animal and vegetable tissues.

Some living structures, such as succulent fruit, jelly-fish, etc., consist almost entirely of water. In all these natural conditions water is impure. Rain contains ammonium salts, especially nitrates and nitrites, and when collected in populated places is often decidedly impure; rivers and springs are contaminated generally with sulphates, chlorides, carbonates and organic substances. The total quantity ordinarily dissolved varies from five to thirty grains to the gallon. When the quantity greatly exceeds this, and especially when peculiar substances, such as iron or sulphur, are present, it constitutes a *mineral water*. Sea-water is very rich in mineral substances, and may be regarded as a mineral water.

The more important varieties of natural waters are—

Alkaline or carbonated waters, containing various carbonates in solution, generally with a quantity of free carbonic acid.

Hard waters, in which a decided amount of magnesium or calcium salts is present. The chemical relations of these waters will be hereafter explained.

Chalybeate waters, containing iron, generally as ferrous carbonate with excess of carbonic acid.

Sulphur waters, containing sulphuretted hydrogen and other sulphides.

Acid waters, containing some of the stronger acids in the free state.

Saline or aperient waters, having large amounts of chlorides and sulphates.

The substances thus dissolved influence the health of those regularly drinking the water, but as to the exact nature of the effects little is known. Springs of chalybeate, alkaline and other mineral waters are resorted to by invalids, but how much of the apparent beneficial action is due to the water, and how much to change of scene, diet and other conditions, is a mooted point. With regard to ordinary drinking water, its composition may vary within wide limits without appreci-

able effect upon those who drink it. According to the generally accepted method, the important points to be determined in the analysis of water are—total amount of dissolved matter, degree of hardness, amount of chlorine, nitrates and nitrites, and of organic matter. The difficult problem is the determination of the organic matter. By the term are meant the animal and vegetable substances that get in with the waste, drainage and sewage which all flowing water receives. These substances occur only in minute quantities even in impure water, but appear to be capable of doing much harm.

Purification of Water. Filtration through porous materials, paper, charcoal, sand, etc., will remove suspended impurities. Animal charcoal and spongy iron remove some of the dissolved matters, but distillation is the only method of preparing absolutely pure water.
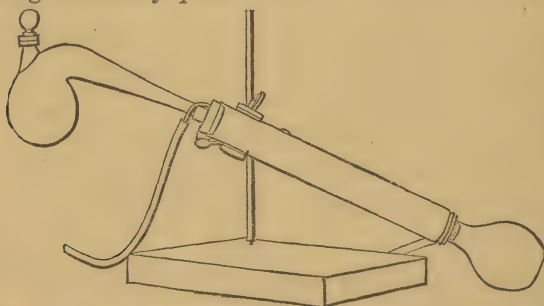
Gen. Chem. Rel. A very large number of substances form, with water, definite chemical compounds, in some of which the water unites without changing; in others the molecule H_2O seems to be broken up. Of the first kind of combination instances are seen in common crystals. Copper sulphate, for instance, has the formula $CuSO_4$, but the blue crystals sold as copper sulphate contain, in addition to this formula, a large amount of water. This water is derived from that in which the copper sulphate was dissolved. It is an essential part of the crystal, for if we drive out all of it the mass is converted into a white powder. Chemical an-

alysis shows that the composition of the blue substance is $CuSO_4 + 5H_2O$. Water that is in this way part of a molecule and essential to a crystalline form is called WATER OF CRYSTALLIZATION. Substances that do not contain it in such a state of combination are said to be ANHYDROUS. Some common salts form crystals containing large amounts of water; sodium carbonate in its commercial form, $Na_2CO_3 + 10H_2O$, contains over fifty per cent. water. Water of crystallization is usually easily driven out by heat.

The second form of the chemical action of water is where it is apparently decomposed, its two elements associating themselves independently with the elements of the other body. If we mix water with about three times its weight of common quicklime, a violent action, attended with production of much heat, ensues, and a dry powder results, from which no appreciable amount of water can be expelled except by a red heat. If this compound contained water in an unchanged form, a moderate heat would drive it all out; hence, chemists have regarded the compound not as CaO,H_2O, but as CaH_2O_2, which is an entirely new body, containing really neither quicklime nor water. A considerable number of oxides are capable of uniting thus with water and forming bodies known as HYDRATES.

Perhaps the most scientific view of these hydrates is to regard the water as acting the part of an acid (it might indeed be called hydric acid), and just as $CaSO_4$ is calcium sulphate, CaH_2O_2 will be calcium hydrate. Water, in fact, might be written HHO, the first H being replaceable by an element, according to the usual law of atomicity. A portion of the hydrogen must always remain, or the body would become an oxide. Thus, if potassium were to act upon water according to the reaction $K_2 + H_2O = K_2O + H_2$, or calcium were to act according to the reaction
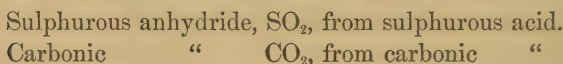
$$Ca + H_2O = CaO + H_2,$$

the bodies produced would not be hydrates, but oxides.

The oxides which, by addition of water, produce hydrates are called ANHYDRIDES. By subtracting H_2O from any hydrate we may reproduce the corresponding anhydride. Many of the common acids may in this way furnish anhydrides, some of which are interesting bodies:
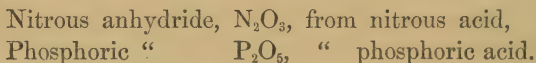
Sulphuric acid.   Sulphuric anhydride.
H_2SO_4 $— H_2O$ $= SO_3$
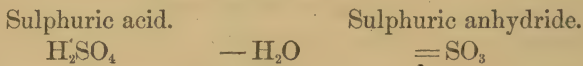
In the same manner the student may deduce

Sulphurous anhydride, SO_2, from sulphurous acid.
Carbonic   " CO_2, from carbonic   "

If the acid contains but one atom of hydrogen, we must, of course, double the formula before subtracting. Hence

Nitric acid.   Nitric anhydride.
$2HNO_3$ $— H_2O$ $= N_2O_5$

Similarly we deduce

Nitrous anhydride, N_2O_3, from nitrous acid,
Phosphoric " P_2O_5, " phosphoric acid.

We may proceed in the same way with other hydrates:

Calcium hydrate.   Calcium anhydride.
CaH_2O_2 $— H_2O$ $= CaO$

As before, when the hydrate contains but one atom of hydrogen we double the formula:

Potassium hydrate.   Potassium anhydride.
$2KHO$ $— H_2O$ $= K_2O$

The term anhydride generally refers to those bodies which yield acids by addition of water. Those which yield hydrates capable of neutralizing acids are generally called bases.

By exposing water to the action of nascent oxygen it can be made to take up an additional atom, and becomes H_2O_2.

Hydrogen Dioxide, H_2O_2. This body, sometimes called oxygenated water, is prepared by liberating oxygen in the

presence of water, as when barium dioxide is dissolved in dilute hydrochloric acid:

$$BaO_2 + 2HCl + H_2O = BaCl_2 + H_2O + H_2O_2.$$

A dilute solution may be easily obtained in this way, but the concentrated liquid is very difficult to prepare. It is a colorless, transparent, oily liquid, nearly one-half heavier than water; it is without odor, has a bitter taste, blisters the skin and bleaches organic colors. It is decomposed by heat and by many chemical substances, often explosively. It dissolves in ether, and the solution has been used in medicine and for bleaching the hair.

POTASSIUM, K, 39.

Sources and Preparation. Potassium occurs in many rocks and soils; from these it is absorbed by land-plants, in the ashes of which potassium carbonate is found. Large deposits of nitrate and chloride also occur. It is best prepared by heating the carbonate with charcoal. Discovered by Davy in 1807.

Properties and Gen. Chem. Rel. Potassium is quite soft, and the freshly-cut surface has a silver lustre, but it quickly tarnishes in the air. It decomposes water rapidly, the escaping hydrogen being so highly heated as to take fire, burning with a purple flame due to the presence of potassium. Specific gravity, 0.865. It is highly electropositive, and forms several oxides, only one of which is important.

Potassa, K_2O, is obtained by oxidizing potassium in dry air. It is generally seen as hydrate.

Potassium Hydrate, *Caustic Potassa,* KHO, is made by boiling potassium carbonate with slaked lime.

$$CaH_2O_2 + K_2CO_3 = 2KHO + CaCO_3.$$

7 *

The solution is filtered from the insoluble $CaCO_3$, evaporated to dryness, the residue fused and cast in sticks. Caustic potassa is a white solid, very soluble in water, powerfully alkaline and corrosive.

Potassium Carbonate, K_2CO_3, *Salt of Tartar.* This is found in the ashes of land-plants, being produced by the action of heat upon the compounds of potassium with organic acids. The ashes are treated with water; the solution thus obtained yields, on evaporation, the impure carbonate termed pearl-ash. Pure potassium carbonate is white, soluble in water, alkaline and moderately corrosive.

Acid Potassium Carbonate, $KHCO_3$, *Salœratus,* is prepared by adding carbonic acid to the normal carbonate. It is a white crystalline body, soluble in water, and is used in effervescing mixtures, but acid sodium carbonate has of late years substituted it to a great extent. It is often called *bicarbonate.*

Potassium Sulphate, K_2SO_4. This is a residue of some manufacturing operations. It forms hard, colorless, six-sided crystals, which are not very soluble in cold water.

Acid Potassium Sulphate, $KHSO_4$, is also a by-product in certain operations. It is sour and strongly acid to test-paper. It is often called bisulphate, and is used as a substitute and adulterant for cream of tartar.

Potassium Nitrate, KNO_3, *Nitre, Saltpetre,* is found on the surface of the soil in India, and is prepared artificially by allowing nitrogenous matter to decay in the presence of wood-ashes (containing K_2CO_3) and lime, and in a full supply of air. It appears that ammonia is first formed and then oxidized. The crude nitre, extracted by water, must be carefully purified, when it appears in large crystals, soluble in water. It melts below redness, and when further heated decomposes, giving off oxygen and nitrogen, and leaving K_2O. Potassium nitrate is used in gunpowder and fireworks as a source of oxygen. Gunpowder consists of about

75 parts nitre, 15 parts charcoal and 10 parts sulphur. The reaction is approximately

$$4KNO_3 + C_4 + S = K_2CO_3 + K_2SO_4 + N_4 + 2CO_2 + CO.$$

The N, CO_2 and CO occupy at the moment of explosion about 1200 times the bulk of the powder, and the explosive action of gunpowder is due to this sudden expansion in volume.

Potassium Chlorate, $KClO_3$. The method of manufacture is given on page 63. The salt crystallizes in flat, tabular crystals. It melts below a red heat; at a little higher temperature gives off all its oxygen, leaving KCl. It is not very soluble in cold water. It is used largely as a source of oxygen, also in matches and fireworks, and as a medicine.

Exp. Some crystals of potassium chlorate and a few pieces of phosphorus are put into a wineglass nearly filled with water, and sulphuric acid poured directly upon them by means of a funnel-tube. The chlorate decomposes, furnishing various oxides of chlorine, which cause the phosphorus to burn brilliantly under the water.

Exp. Powdered sugar mixed with about three times its weight of potassium chlorate will burn when touched with a drop of sulphuric acid.

Exp. A mixture of two parts potassium chlorate with one part of potassium ferrocyanide and one part sugar makes *white* gunpowder, which explodes very easily and very violently. Only a small quantity should be made at once, and the ingredients should be powdered separately and mixed gently.

Potassium Chloride exists in sea-water and in the saline deposits at Stassfurth, Germany. It resembles common salt. It forms an insoluble double salt with platinic chloride, $2KCl + PtCl_4$.

Potassium Bromide is made according to the method given on page 64. It forms cubical crystals soluble in water.

Potassium Iodide is prepared like the bromide, which it closely resembles, but is rather a finer white. It is easily soluble in water.

KBr and KI are often made by first preparing the corresponding iron compound, and decomposing it with potassium carbonate.

Tests. Potassium compounds are mostly soluble in water. A few, however, are so slightly soluble as to afford us serviceable tests.

1. Platinum chloride produces a yellow crystalline precipitate of potasso-platinum chloride, $2KCl + PtCl_4$.

2. Tartaric acid gives a white crystalline precipitate of acid potassium tartrate, $KHC_4H_4O_6$.

Brisk stirring with a glass rod promotes the formation of both precipitates.

3. Potassium compounds give to flame a color which is a mixture of red and violet.

SODIUM, Na, 23.

Sources. Common salt, NaCl, is the principal source. Sodium compounds are widely distributed, occurring even in common dust. The ashes of sea-plants contain sodium carbonate. Sodium was discovered by Davy in 1807.

Preparation and Properties. It is prepared in a manner similar to potassium, which it closely resembles, but is a little heavier and not so easily oxidized. Its chemical relations and the properties of its compounds are also much like those of potassium.

Sodium Hydrate, NaHO, *Caustic Soda*, is prepared by a process similar to that used for caustic potash, using Na_2CO_3 instead of K_2CO_3. Caustic soda is usually sold in cylindrical sticks. It is soluble in water, and is very strongly alkaline and corrosive.

Sodium Carbonate, Na_2CO_3, *Sal Soda*, was formerly ob-

tained from the ashes of sea-plants, but is now made principally by the action of chalk and charcoal upon sodium sulphate. It forms large crystals, having the composition $Na_2CO_3 + 10H_2O$, very soluble in water, and often called *washing soda*. On exposure to air these crystals *effloresce*— that is, lose water and fall to a white powder.

Acid Sodium Carbonate, $NaHCO_3$, *Baking Soda*. This body is produced in the same manner as acid potassium carbonate, which it closely resembles. It is now much used in effervescing mixtures like Seidlitz powders and the common baking-powders, which latter are usually a mixture of cream of tartar and baking soda. Alum and acid potassium sulphate are often used in the inferior grades as a substitute for the cream of tartar. The action of the powder is due to the sudden evolution of a large volume of carbon dioxide,

$$NaHCO_3 + KHC_4H_4O_6 = NaKC_4H_4O_6 + H_2O + CO_2.$$

Sodium Sulphate, *Glauber's Salt*, Na_2SO_4, is a by-product in the manufacture of nitric and muriatic acids. It forms large clear crystals which contain ten molecules of water of crystallization. They effloresce in dry air, and are remarkable for being more soluble in water at 93° F. (34° C.) than at any other temperature. The principal use of sodium sulphate is as a source of sodium carbonate.

Sodium Nitrate, $NaNO_3$, is found in large beds in northern Chili, and termed Chili saltpetre. It is used as a manure, and also in the preparation of nitric acid. It is not used in gunpowder, on account of its tendency to absorb water.

Sodium Chloride, *Common Salt*, $NaCl$, is too well known to need description. It occurs in thick beds in various parts of the world, and is also prepared from sea-water by evaporation or freezing, and from certain brine-springs by evaporation. It dissolves in about the same amount in hot and cold water.

Sodium Phosphates. The only important form is disodium

acid phosphate, Na_2HPO_4, which is used in medicine and also as a test for magnesium.

Sodium Anhydroborate, $2NaBO_2 + B_2O_3$, commonly called sodium biborate or borax, is found in certain lakes in Thibet and in California. It is also made by melting sodium carbonate with boric acid. It forms hard crystals, which dissolve in about twelve times their weight of water and form an alkaline solution. Borax is much used as a solvent for metallic oxides, especially in blowpipe analysis. It is used for cleaning metals in soldering.

Sodium Silicate, made by fusing sand or pulverized quartz with an excess of sodium carbonate, constitutes *soluble glass,* which dissolves in boiling water. It is used as a cement and in soaps.

Sodium Thiosulphate, $Na_2S_2O_3$, much used in photography under the name *hyposulphite.* Its solution possesses the power of dissolving most of the salts of silver, except Ag_2S, which are insoluble in water.

Sodium Sulphite, Na_2SO_3, is used as a substitute for sulphurous acid in preventing fermentation.

Tests. The only convenient test for sodium is the strong yellow color that its compounds give to flame.

Lithium, Li, 7. Discovered by Arfvedson in 1817. It is found in many substances, but only in small quantity. Its principal sources are some rather rare minerals. Lithium resembles potassium. Sp. Gr. 0.593. Its salts resemble those of potassium and sodium, but lithium carbonate is but sparingly soluble in water. Lithium compounds have been used in gout and similar chronic diseases; and, as many spring waters contain traces of lithium, the medicinal action of such waters has been supposed to be due to it, and much nonsense and quackery have been developed in connection with the analysis of them.

Tests. Lithium imparts a crimson color to flame, which is the most convenient and delicate method of detecting its compounds.

Cæsium, Cs, 133, and **Rubidium,** Rb, 85.4, were discovered by Bunsen and Kirchoff in 1860 by the spectroscope. They exist only in small quantity in some mineral waters and in a few plants. They are strongly positive and closely resemble potassium. Cæsium gives a blue color to flame; rubidium, a dark-red color.

SILVER, Ag, 108.

Sources. Silver occurs native—that is, in the free state— in moderate abundance, also as sulphide, chloride and other forms. It is often present in small amounts in lead ores. Silver was known to the ancients.

Preparation. Silver is easily reduced to the metallic state, most of its compounds being decomposed by heat. When existing in small quantities in ores, it may be taken out by agitating the powdered material with mercury, which dissolves the silver (amalgamation), and this amalgam, being drawn off and distilled, leaves the metallic silver. When compounds like sulphides or chlorides are reduced, it is some- times necessary to add iron scraps to liberate the silver before adding the mercury.

Properties. · Silver is white and highly lustrous, easily worked into plates and wire, and the best conductor of heat and electricity known. Specific gravity, 10.5. It resists the action of oxygen and of caustic alkalies, but is attacked by sulphur and sulphides and by nitric acid. Solutions of silver salts are decomposed by heat, light and electricity, and by many forms of organic matter, especially when mixed with alkalies. The sensitiveness of its salts to light is the basis of

photography. It melts at 1681° F. (916° C.). For coinage it is usually alloyed with copper. In some of its properties— *e. g.* its high specific gravity, resistance to the action of the air and the slight solubility of its oxide and carbonate—it is related to lead and copper.

Silver Oxide, Ag_2O, cannot be formed directly, as, although oxygen is absorbed by melted silver, no combination is formed on cooling. It is a black powder, usually made by heating *silver hydrate*, $AgHO$, which latter is produced when silver nitrate is mixed with an alkali.

$$AgNO_3 + NaHO = NaNO_3 + AgHO.$$

Silver Sulphate, Ag_2SO_4, is sometimes used in analysis.

Silver Nitrate, $AgNO_3$, *Lunar Caustic*, is easily made by dissolving the metal in nitric acid. The reaction is similar to that with copper.

$$Ag_3 + 4HNO_3 = 3AgNO_3 + 2H_2O + NO.$$
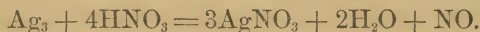
Silver nitrate forms colorless crystals, very soluble in water, and, when mixed with organic matter, blackened by light. It fuses at 426° F. (219° C.), and is often cast in sticks for use as a caustic. The property of forming a black, difficultly soluble precipitate with organic matter is utilized in the manufacture of hair-dyes and marking-ink.

Silver Chloride, $AgCl$, is found as a mineral, and is easily formed artificially by adding any chloride to silver nitrate.

$$NaCl + AgNO_3 = AgCl + NaNO_3.$$

It forms a heavy white precipitate like curdled milk, turning violet in the light, especially if organic matter be present, forming a subchloride, Ag_2Cl. It is insoluble in most acids, but dissolves freely in ammonia and in sodium thiosulphate.

A number of other silver compounds have been described, but have but little importance.

Silver Salts in Photography.—The action of light upon silver

compounds is mostly an operation of reduction, sometimes to the metallic condition. The presence of organic matter aids the change. In some of the compounds the effect of the light is not attended with any change of color.

Oxygen Group. This group includes oxygen, sulphur, selenium and tellurium. They are electro-negative dyads, and have a wide range of affinity, combining with many bodies in several proportions. With many substances they unite in small proportion to form bases, and in large proportion to form anhydrides.

OXYGEN, O, 16.

Sources. Oxygen exists in water, air, all animal and vegetable tissues and in the great majority of minerals. It constitutes over half the matter composing the earth. It was discovered by Scheele and Priestley in 1774. The name means " producer of acids."

Preparation. 1. The oxides of mercury and of some other elements are decomposed by heating. This method is of little practical value, but is interesting, as it was the means of the discovery of the gas, $HgO = Hg + O$.

2. Certain compounds of manganese and barium when heated in a current of air absorb oxygen, and give it out in a current of steam. By alternating these two currents large quantities of oxygen may be obtained. Such processes are not suitable for laboratory-work.

3. The chlorates, nitrates and some rarer salts are decomposed by heat, giving off large quantities of oxygen, but not always quite pure. Potassium chlorate is by far the most suitable. Used alone, it requires a high temperature, but when mixed with about one-quarter of its weight of manganese dioxide a heat of about 500° F. (260° C.) is sufficient.

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The exact manner in which manganese dioxide acts has not been explained.

Exp. Mix thoroughly 4 parts potassium chlorate with 1 part manganese dioxide, and heat the mixture in any suitable vessel. One ounce of potassium chlorate yields nearly two gallons of the gas. It may be collected in a glass bottle held over the end of the tube, this bottle having been previously filled with water and then inverted into the bowl, after closing the mouth of it with a card or glass plate. No water will escape until bubbles from the tube are passed into it, which, on account of their lightness, ascend and displace the water. When the water is all out, remove the bottle, and place it mouth downward in a saucer of water, replacing it with another bottle previously filled with water, and repeat this process until the evolution of gas ceases. The first bubbles that pass over consist of air contained in the test-tube, and the pure gas quickly succeeds.

The reaction in this experiment concerns the potassium chlorate only, which is simply decomposed:

$$KClO_3 = KCl + O_3.$$

The arrangement depicted in the cut is called the pneumatic trough, and is extensively used for collecting gases. Water is the liquid generally employed, but in some cases mercury is used. As long as the vessels in which the gas is collected are kept inverted and their mouths below the surface of the water (or mercury), no gas can escape or air get in.

Properties. Pure oxygen is colorless, odorless and tasteless; it is one-tenth heavier than air, one litre weighing 1.43 grm. It is continually being absorbed by living animals in

the process of respiration, to which function it is essential. It is also consumed in all ordinary combustion. The pure gas causes considerable excitement of the vital functions of animals and greatly increased action in ordinary flames, but does not, as a rule, produce spontaneous combustion.

Exp. A taper or splinter of wood lighted and blown out in such a way as to leave a glowing coal is instantly re-lighted, with a slight explosion, on putting it into the gas. The experiment may be repeated several times.

Exp. A bit of bark or knot of charcoal ignited at one point and plunged into the gas burns brilliantly, producing a colorless gas which combines with water and forms an acid.

Exp. Sulphur burns in oxygen with moderate brilliancy, producing a colorless, highly irritating gas, which in combination with water produces an acid. A picture painted with a solution of quinine sulphate is almost invisible by ordinary light, but becomes visible when illuminated by burning sulphur; the explanation of this effect belongs to physics; it is called fluorescence, and is shown by many substances.

Exp. Phosphorus burns in the gas with great brilliancy, producing dense white clouds, which rapidly absorb water and produce a powerful acid.

It will be noticed that in the foregoing experiments acids have been produced by the combustion. Lavoisier supposed that oxygen was necessary to the production of an acid; the present name of the gas means " acid-producer," and expresses this view. According to the opinion of modern chemists, hydrogen is the essential element of acids, and the following experiments will show that the combination of oxygen may produce bodies very different from acids :

Exp. Magnesium ribbon ignited and plunged into oxygen burns rapidly and brightly, producing a bulky white powder, insoluble in water and of alkaline properties.

Exp. One end of a thin steel ribbon (watch-spring) is wrapped with a few turns of cotton thread and then dipped into melted sulphur or wax. This end being lighted, the ribbon is put into oxygen; the metal quickly takes fire, produces an abundance of sparks, and from time to time drops hot

globules to the bottom of the jar, which often fuse themselves into the glass. A layer of three or four inches of water or an inch of sand will generally prevent the breaking of the jar. The product is black, insoluble in water and destitute of acid properties.

Sodium, potassium and zinc turnings also burn in oxygen with more or less brilliancy.

The explanation of the production of acids in the first set of experiments is that the water, which is always present in these cases, forms with the product of combustion a new substance. The combinations of oxygen are called oxides; in the case of the burning sulphur the resulting gas has the formula SO_2, sulphur dioxide, and if the materials were perfectly dry no acid would be formed. In the presence of water we have the reaction $H_2O + SO_2 = H_2SO_3$, sulphurous acid. Similarly, the charcoal produces CO_2, which, uniting with water, produces H_2CO_3, carbonic acid. The same principle applies to the phosphorus experiment.

On the other hand, the metals manganese, sodium, etc. form oxides which, so far from being acids, are really powerful neutralizers of acids and are said to be alkaline or basic. They combine with water, producing compounds which retain the alkaline character. In this way we have $MgO + H_2O = MgH_2O_2$, $Na_2O + H_2O = Na_2H_2O_2$. Further explanations will be found in the section on Water.

Gen. Chem. Rel. Oxygen combines with every other element except fluorine, and with many in several proportions. The chemical relations of these oxides are dependent in part upon the number of oxygen atoms present. The oxides of manganese may be taken as examples:

MnO,	Powerful base.
Mn_2O_3,	Weak base.
MnO_2,	Indifferent.
MnO_3,	Forming an acid (anhydride).

These instances illustrate the general law that small proportions of oxygen tend to produce bases, high proportions anhy-

drides or acid-forming oxides, and intermediate proportions bodies of uncertain or indifferent character. Some elements are apparently incapable of yielding bases. These form with oxygen, in low proportions, neutral oxides. This is shown by the nitrogen series:

N_2O,	Indifferent.
NO,	Indifferent.
N_2O_3,	Acid-forming.
NO_2,	Doubtful.
N_2O_5,	Acid-forming.

Oxygen is a dyad, and is generally considered the most electro-negative element, but under some circumstances chlorine seems to be superior in this respect. It is slightly soluble in water, and upon this fact depends the existence of most forms of aquatic life.

Substances which take away oxygen from its combinations are called *reducing agents;* those which add oxygen, *oxidizing agents.*

Tests. Besides the power of relighting a taper, free oxygen may be recognized by its turning brown a mixture of caustic soda and pyrogallin, and by converting colorless nitric oxide into red peroxide.

Ozone. Oxygen is susceptible of a modification of some of its properties without chemical change. The study of this subject was begun in 1840 by Schönbein, and very many researches have since been made, without, however, completely explaining the condition. The modified oxygen is called " ozone," from a Greek word meaning " to smell," on account of its marked odor.

Ozone may be prepared in several ways:

1st. By a succession of electrical sparks through air or oxygen. Its peculiar odor is observed when an electrical machine is put into active operation.

8 *

2d. By the slow oxidation of phosphorus and of turpentine and other essential oils.

3d. By the decomposition of water by the galvanic current.

4th. By the action of acids upon certain bodies rich in oxygen.

By all these methods only a small proportion of the oxygen is converted into ozone.

Exp. Place a few crystals of potassium permanganate in a wide-mouthed bottle and add a few drops of sulphuric acid. A peculiar odor is noticed; the evolved gas will tarnish mercury and silver, and turn blue a piece of paper soaked in a solution of potassium iodide and starch. This last effect is due to the setting free of the iodine by the superior affinity of the ozone for the potassium. Organic matter is also acted upon ·powerfully.

Ozone is heavier than oxygen, is soluble in water, and is converted into common oxygen by heat and by contact with some oxides. It is generally present in the atmosphere, especially in open country-places. It is considered an important natural disinfectant.

Another modification of oxygen, called antozone, has been supposed to exist, but this is considered by many chemists to be hydrogen dioxide, H_2O_2.

SULPHUR, S, 32.

Sources. Sulphur occurs native—*i. e.* in the free state—in volcanic regions, also in combination, forming sulphides and sulphates, and in animal and vegetable structures.

Preparation. Commercial sulphur is prepared by melting or distilling the native sulphur or some of the sulphides. It presents itself in two forms: *roll sulphur or brimstone*, made by casting the melted sulphur in moulds; and *flowers of sulphur*, made by condensing the distilled sulphur in a cool

chamber. *Lac sulphuris*, or milk of sulphur, is a finely-divided medicinal form, obtained by dissolving common sulphur in milk of lime, and precipitating by an acid.

Properties. Sulphur assumes several allotropic forms, varying especially in color and solubility; it is dimorphous —*i. e.* crystallizes in two forms, octahedral and prismatic. Ordinarily, sulphur is brittle, yellow and soluble in carbon disulphide, but by being suddenly cooled from near its boiling-point it becomes plastic, dark-colored and insoluble. All varieties are insoluble in water, highly combustible, fusible at about 250° F. (121° C.) and boiling at 836° F. (447° C.). It is a non-conductor of electricity, and becomes highly electrical by friction.

Exp. A small quantity of sulphur is placed in a flask and heated slowly. It melts to a thin, amber-colored liquid. On continuing the heat the liquid gradually becomes thick, and at about 450° F. (232° C.) it is so tenacious that it can scarcely be poured out of the vessel. Heated still further, it becomes thin, and finally boils. Just before the sulphur begins to boil, pour it into cold water, when it will form dark brown, semi-elastic masses. If the sulphur remaining in the flask be heated to the boiling-point, a dark red vapor is produced, in which certain substances—*e. g.* Dutch leaf —burn easily.

Exp. Dissolve some sulphur in a small quantity of carbon disulphide, and allow the solution to evaporate in the open air. Transparent yellow octahedra will be deposited.

Exp. A few ounces of sulphur are fused in a Hessian crucible, and then allowed to cool until a film of solid forms on the surface. Break a small hole through this and pour out the liquid contents. On breaking the crucible a mass of prismatic crystals will be found.

Gen. Chem. Rel. Sulphur forms an important group of compounds—the sulphides. In these it is electro-negative and usually diatomic. With the members of its own group it
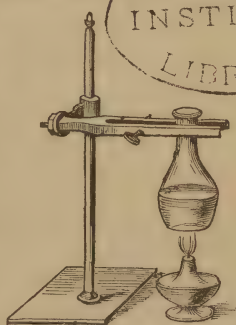
combines in several proportions, showing atomicities of two, four and six, and perhaps even higher. In association with oxygen and chlorine it is regarded as electro-positive. In general its compounds are analogous in composition to those of oxygen, and as many oxides act as bases toward the ordinary acids, so the corresponding sulphides act as bases toward what are called the sulphur acids. Thus we have

$$K_2O + CO_2 = K_2CO_3 \quad \text{Potassium carbonate.}$$
$$K_2S + CS_2 = K_2CS_3 \quad \text{``} \qquad \text{sulphocarbonate.}$$
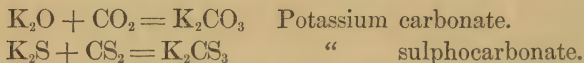
In such compounds the sulphur is substituted for the oxygen, atom for atom, and the name is formed by affixing the syllable "sulph" to the name of the acid. A few of these sulphur salts are very important.

Sulphur has many important uses. It is employed in medicine as an alterative and externally for skin diseases. It is used in the arts for vulcanizing caoutchouc and in the manufacture of gunpowder. Match-sticks are tipped with it to make the friction composition ignite the wood more surely.

Sulphur forms two compounds with hydrogen:

H_2S Hydrogen sulphide or sulphuretted hydrogen.
H_2S_2 Hydrogen disulphide.
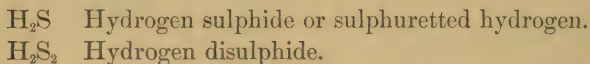
Hydrogen Sulphide, H_2S. Discovered by Scheele in 1777. This substance is a gas. It exists in solution in some spring waters, also in the emanations from volcanoes and decomposing animal and vegetable matters. It is sometimes produced by the action of organic matter upon sulphates. Calcium sulphate, $CaSO_4$, by losing its oxygen becomes CaS, and this, by the action of carbonic acid, yields the gas, $CaS + H_2CO_3 = CaCO_3 + H_2S$. Hydrogen sulphide is much used as a test, and is frequently made in the laboratory by acting upon sulphides with strong acids. Ferrous sulphide and sulphuric acid are much used. $FeS + H_2SO_4 = FeSO_4 + H_2S$.

Exp. Introduce into a gas-bottle some ferrous sulphide in small fragments, and pour over it some dilute sulphuric or hydrochloric acid.

The hydrogen sulphide is liberated freely. As it is only slightly heavier than air and rather soluble in water, it is difficult to collect either by displacement or over water.

Properties. Hydrogen sulphide is a colorless gas, condensible by moderate cold and pressure. It has a strong odor like rotten eggs, is easily combustible, burning with a pale blue flame and producing sulphurous acid, H_2SO_3. Water at ordinary temperature dissolves about three volumes, acquiring the odor and chemical properties of the gas. The important property of hydrogen sulphide is its power of precipitating many elements as sulphides. These precipitates being generally distinct in color and highly insoluble, their production is not only a test for the presence of such bodies, but also a means of separating them from solution.

Exp. As illustrations, solutions of copper sulphate, mercuric chloride, tartar emetic and zinc sulphate may be precipitated either by a current of the gas or by its aqueous solution. The zinc sulphate requires the addition of a few drops of ammonia.

When pure hydrogen sulphide is required it is generally obtained by heating a mixture of hydrochloric acid and antimony sulphide. $Sb_2S_3 + 6HCl = 2SbCl_3 + 3H_2S$.

Hydrogen sulphide is a powerful reducing agent; solution of potassium permanganate is rapidly decolorized by it. A mixture of potassium bichromate and hydrochloric acid is changed into green chromic chloride.

Tests. Hydrogen sulphide is recognized by its smell and by its blackening paper soaked in lead acetate solution. As the gas is a frequent product of decomposition, the leakage of sewer-gas may sometimes be detected by this test.

Hydrogen Disulphide, H_2S_2. This substance may be prepared by boiling together lime and sulphur, forming calcium disulphide, and pouring the solution into dilute hydrochloric

acid. The hydrogen disulphide separates as a yellow oily liquid of a disagreeable odor. It decomposes easily.

Compounds of Sulphur with Oxygen.—Sulphur forms with oxygen a number of acid-forming oxides or anhydrides, most of which are known only in the hydrated condition— that is, as acids:

Anhydride.		Acid.	Name.
SO		H_2SO_2	Hyposulphurous.
SO_2		H_2SO_3	Sulphurous.
SO_3		H_2SO_4	Sulphuric.
S_2O_2		$H_2S_2O_3$	Thiosulphuric.
S_2O_5		$H_2S_2O_6$	Dithionic.
S_3O_5	not yet obtained.	$H_2S_3O_6$	Trithionic.
S_4O_5		$H_2S_4O_6$	Tetrathionic.
S_5O_5		$H_2S_5O_6$	Pentathionic.

The first member of this series is the true hyposulphurous acid, the commercial hyposulphites being really salts of thiosulphuric acid, hence properly called thiosulphates. Most of the members of the above series are of limited interest.

Sulphur Dioxide, Sulphurous Anhydride, SO_2. This substance is found in the emanations from volcanoes and frequently in the air of towns, being derived in the latter case from the sulphur in coal and coal gas. It is the usual product of the burning of sulphur or the sulphides in air or in oxygen, and is generally so made on the large scale. For laboratory operations it is most conveniently obtained by deoxidizing sulphuric acid. The experiment requires strong acid and considerable heat. The substances which produce the best results are not soluble in the dilute acid. Copper, mercury, charcoal, silver, sulphur and other bodies may be used; the first mentioned is the best for a small experiment. The sulphur dioxide which exists in the atmosphere of towns is slowly oxidized to sulphuric acid.

Exp. Strong sulphuric acid and some copper turnings or twisted copper wire are put into a flask and heated carefully. The gas does not come off until the temperature gets rather high. It may be collected by downward displacement, as shown under chlorine.

The pneumatic trough cannot be used, as the gas is very soluble in water. To show this properly, the end of the delivery-tube may be dipped in some water in a beaker, when it will be found that most of the bubbles will be taken up by the liquid. This absorption must be carefully watched, or the water may be drawn back into the hot sulphuric acid.

The reaction in the preparation of sulphurous acid is rather difficult to understand. When strong positives, like zinc or magnesium, are put into acids, the usual result is the expulsion of hydrogen. With zinc and sulphuric acid we get $Zn + H_2SO_4 = ZnSO_4 + H_2$. With bodies of less positive character a portion of the oxygen of the acid is removed, and water, sulphurous anhydride and a sulphate are formed. Copper gives

$$Cu + 2H_2SO_4 = CuSO_4 + 2H_2O + SO_2.$$

Mercury and silver also give similar effects. Carbon and sulphur also deoxidize the acid, but no sulphates are formed, as these bodies cannot form salts.

Carbon and sulphur give, respectively, the following reactions :

$$C + 2H_2SO_4 = 2SO_2 + 2H_2O + CO_2.$$
$$S + 2H_2SO_4 = 3SO_2 + 2H_2O.$$
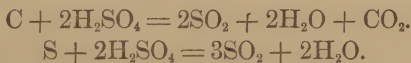
The carbon reaction is the most economical, but the admixed

CO_2 is objectionable. Sulphur dioxide may be made by heating an intimate mixture of manganese dioxide and powdered sulphur.

$$MnO_2 + S_2 = SO_2 + MnS.$$
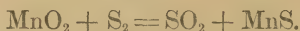
Properties. It is a colorless gas, of the well-known irritating odor of burning matches. It can be condensed to a colorless liquid by a cold $0°$ F. ($—18°$ C.). A mixture of snow and salt answers quite well as a means of obtaining sufficient cold. The liquid so obtained is sulphur dioxide, SO_2, and not sulphurous acid. It boils at $14°$ F. ($—10°$ C.) and freezes at $—105°$ F. ($—76°$ C.). Sulphur dioxide passed into water forms sulphurous acid, $H_2O + SO_2 = H_2SO_3$, which remains in solution, giving the liquid all the common properties of an acid; vegetable colors are first reddened, then bleached. The solution is also a powerful reducing agent, and is much used for that purpose.

Although, theoretically, we make a distinction between the anhydride and the acid, yet, practically, we disregard this distinction, and for most experiments may use either the gas or the solution in water.

The principal properties of this body may be easily shown by the following experiments:

Dip a lighted taper in a jar in which the gas has been collected by displacement. The flame is immediately extinguished.

Suspend a delicately colored flower, somewhat moist, in a jar of the gas. The colors bleached by this agent are not entirely destroyed, and by exposure to the action of weak ammonia are often restored, with curious modifications.

Add some solution of sulphurous acid to solutions of

(*a*) potassium bichromate mixed with hydrochloric acid;

(*b*) potassium permanganate.

(*a*) will be changed to green chromic chloride;

(*b*) will be wholly decolorized. These effects are due to the reducing action of the sulphur dioxide, which thus becomes converted into sulphuric acid.

The anhydride, free acid and its salts are antiseptic agents —that is, prevent putrefaction and fermentation, especially the latter. For this reason the vapors of burning sulphur are extensively used for fumigating ships and other places in which infectious diseases may exist. Wine- and beer-casks are also purified by sulphur, and the sulphites are added to fermented liquor to prevent further change. The acid is supposed to act by killing the minute living structures which are nearly always developed in decomposing and fermenting substances.

Sulphur dioxide and hydrogen sulphide decompose each other according to the following reaction:

$$2H_2S + SO_2 = 2H_2O + S_3.$$

The experiment may be easily performed by passing currents of the two gases into a small bottle.

Gen. Chem. Rel. The salts of sulphurous acid are called sulphites; monads, of course, form two compounds, acid and normal. Thus, potassium gives us

Acid potassium sulphite.   Potassium sulphite.
$KHSO_3.$ $K_2SO_3.$

Dyads give but one sulphite. From calcium we have only $CaSO_3$.

Tests. The odor and bleaching power of the free acid and the reducing properties of the sulphites are sufficient means of detection.

Sulphur Trioxide, Sulphuric Anhydride, SO_3. This body is obtained by distilling Nordhausen sulphuric acid (*q. v.*), and by the action of phosphoric anhydride upon common sulphuric acid. The latter reaction is a simple dehydration.

$$H_2SO_4 + P_2O_5 = H_2P_2O_6 (= 2HPO_3) + SO_3.$$

Several other methods of making it are known.

Properties. Sulphuric anhydride is a soft, white, odor-

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less solid in long, silky crystals like asbestos. Exposed to the air, it absorbs water rapidly and becomes converted into sulphuric acid. When dropped into water the energy of combination is so great that a hissing noise is produced. The dry substance is destitute of corrosive properties, and is now sent into commerce in sealed iron boxes for use in certain manufacturing operations. It is occasionally employed as a means of drying gases and for special purposes in chemical research. It has a specific gravity of 1.95, melts at 65° F. (18.3° C.) and boils at 95° F. (35° C.).

Sulphuric Acid, Hydrogen Sulphate, Oil of Vitriol, H_2SO_4. This substance was probably known to Geber in the eighth century, and was certainly known to Basil Valentine in the fifteenth century. It is now made in large quantities in almost all parts of the civilized world, and it has been said by a technologist that the material prosperity of a country may be judged of by the extent of its sulphuric acid manufacture.

Sources. Sulphuric acid occurs free in waters of volcanic and mining districts, and sometimes in the air of towns, being in the latter case derived from the oxidation of sulphurous acid. It also occurs in the saliva of certain animals. The compounds of sulphuric acid (sulphates) are of frequent occurrence. Calcium and barium sulphates are abundant minerals; sodium sulphate occurs in many natural waters.

Preparation. The original method of preparation was the distillation of the sulphates, especially the ferrous sulphate, $FeSO_4$, which was produced by the oxidation of iron pyrites, FeS_2. The process is still used for special purposes, as the acid so produced is more concentrated than the ordinary commercial article. The method by which the common acid is made depends upon the power of one of the oxides of nitrogen to act as a carrier of oxygen from the air to sulphurous acid. The outline of the practical process is as follows: Vapors of nitric and sulphurous acids are mixed with steam and air in a large leaden room, the floor of which is

slightly inclined and covered by a few inches of water. The sulphurous acid is derived either from the burning of raw sulphur or the roasting of pyrites; the nitric acid, from the action of sodium nitrate on sulphuric acid. The chemical changes are somewhat complicated, and are not wholly understood. The nitric acid changes some sulphurous acid to sulphuric, becoming itself converted into nitric oxide, NO, by this action. This NO takes oxygen from the air, and forms NO_2, which oxidizes more sulphurous acid, and is thus again converted into NO, and again acted upon by the air. It will be seen that a small quantity of nitric acid will be sufficient to oxidize large quantities of sulphurous acid. The presence of large excess of water is essential to the reactions, hence steam, or water in a fine spray, is thrown continuously into the room. The above description includes only the principal reactions; the following series of equations show more exactly the changes occurring:

$$SO_2 + H_2N_2O_6 = H_2SO_4 + 2NO_2.$$
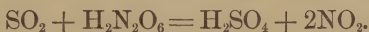
The steam acts upon the NO_2 thus:

$$3NO_2 + H_2O = H_2N_2O_6 + NO.$$

The NO takes O from the air and becomes NO_2, which together with the nitric acid just produced oxidizes more sulphurous acid, reproducing the NO, and the changes go on anew.

The method may easily be shown experimentally:

Exp. A wide-mouthed quart bottle is provided, with a cork perforated for four tubes. To these tubes are attached a flask for generating sulphurous acid, a bottle for evolving nitric oxide (*q. v.*) and a flask for furnishing steam; the fourth tube is left open to the air. Sulphurous acid and nitric oxide are produced, and allowed to flow into the bottle. They combine and produce a white crystalline solid, which quickly disappears when steam is admitted. Some water should also be poured down the open tube. The chemical changes are in part evident to the eye by the change of color which attends the conversion of red NO_2 into colorless NO, and *vice versâ*. After the experiment has lasted ten or fifteen minutes, sufficient dilute acid will have collected in the bottle to respond to the tests given below.

In practice it is found that loss of NO is constantly occur-

ring, and so it is necessary to renew the nitric acid occasionally. The liquid on the floor of the leaden room is drawn off from time to time, and concentrated by boiling in lead pans until it becomes strong enough to attack the lead. The further concentration is conducted in glass or platinum vessels.

Properties. Pure sulphuric acid is a colorless, oily liquid of a specific gravity of 1.848, boiling at about 640° F. (338° C.). It is highly corrosive and poisonous. Exposed to the air, it absorbs water in considerable amounts. When added to water it produces heat, often sufficient to make the water boil, and the dilution of any considerable quantity must be performed with care. A diminution of bulk occurs when the acid and water are mixed. So great is the affinity of sulphuric acid for water that it will decompose many organic substances, extracting the hydrogen and oxygen and leaving the carbon. The carbon so liberated will diffuse through the acid and give it a dark color.

Commercial sulphuric acid is usually more or less brown, or even black, from the carbon set free from particles of dust, straw, etc. which accidentally fall into it. It always contains a small quantity of water—about one molecule to twelve of acid:

$$H_2O + 12H_2SO_4.$$

Its properties, boiling-point, etc. are similar to those of the pure acid.

Nordhausen or Fuming Sulphuric Acid is the original oil of vitriol, so called because it was obtained by the distillation of green vitriol. It is substantially a solution of sulphur trioxide, SO_3, in sulphuric acid. It is denser and even more corrosive than the common acid, and unites with water with great energy. It is used for dissolving indigo and for a few other purposes. When heated, the sulphur trioxide distils off, and the ordinary acid is left.

The properties of sulphuric acid are greatly modified by dilution; its corrosive and charring action may be entirely

removed by adding much water. When such a dilute acid is boiled, it steadily loses water until the original degree of concentration is nearly or quite restored. The following experiments show the more important properties :

Mix in a thin glass vessel equal volumes of strong sulphuric acid and water, and place upon the surface of the liquid a small capsule made of tin or copper foil and containing a small piece of phosphorus. The heat will be sufficient to ignite the phosphorus. If the quantities of water and acid be carefully measured, the amount of condensation may be observed. It will be about one-fifth. The dilute acid so obtained is useful for many experiments, and should be preserved.

The milkiness produced in this mixture is due to the precipitation of lead sulphate formed from the pans in which the acid is concentrated. This body is soluble in the strong, but not in the dilute, acid.

Place a few bits of straw, wood or common organic matter in some strong sulphuric acid. In the course of a few minutes the acid will be discolored by the carbon set free. This discoloration of the acid does not interfere with its ordinary uses.

If any design be traced on white paper with the dilute acid obtained in a former experiment, and the paper then cautiously heated, the acid will slowly become more concentrated, and will finally char the paper completely, but only at the parts which have been touched by the original liquid. This experiment is made use of for the detection of small quantities of the free acid.

The uses of sulphuric acid are very numerous. By its high affinity it is capable of expelling many other acids from combination. Nitric acid is made from nitrates, acetic acid from acetates, by its agency. Its affinity for water makes it a useful drying agent, especially for gases. Many organic bodies are peculiarly modified by treating with dilute sulphuric acid, but these changes will be best understood when described under organic chemistry.

Gen. Chem. Rel. The salts of sulphuric acid are called sulphates. Monads give, of course, two sulphates, acid and normal. Sodium gives us

<div style="text-align:center">

Acid sodium sulphate.   Sodium sulphate.

$NaHSO_4$ Na_2SO_4

</div>

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Dyads give but one sulphate. From barium we get $BaSO_4$, barium sulphate.

Most sulphates, except those of the calcium group, are freely soluble in water.

Tests. The concentrated acid is easily recognized by its oiliness and charring action on organic matter. The dilute acid may be made to produce this charring by evaporation, as described above.

The general test for either the acid or any of its salts is the addition of a solution of some barium compound (barium nitrate, chloride or acetate); a white precipitate insoluble in water and dilute acids is at once formed, even if only a trace of the acid be present. The solution to be tested should be made acid with hydrochloric acid, to prevent carbonates, phosphates, etc. being mistaken for sulphuric acid.

The commercial sulphuric acid contains several impurities. Of these the most important are arsenic and lead, which may be detected and removed by treating the diluted acid with hydrogen sulphide.

Hyposulphurous acid, H_2SO_2, is produced by dissolving zinc in sulphurous acid.

$$H_2SO_3 + Zn = ZnO + H_2SO_2.$$

It is a powerful bleaching and reducing agent, and decomposes quickly.

Thiosulphuric acid, $H_2S_2O_3$, commonly but erroneously called hyposulphurous acid, has not been obtained in the free state. Calcium thiosulphate is formed in several manufacturing operations, and sodium salt is much used in photography. The thiosulphates are powerful reducing agents.

A sesquioxide of sulphur, S_2O_3, has been described, but is not important.

The remainder of the sulphur acids have no particular interest, and will not be described.

Compounds of Chlorine with Sulphur.—These have considerable theoretical and some practical interest, but will be only briefly mentioned.

S_2Cl_2 Sulphur chloride.
SCl_2 "   dichloride.
SCl_4 "   tetrachloride (doubtful).

The first two are liquids. S_2Cl_2 is used as an agent in vulcanizing rubber.

SELENIUM, Se, 79.5.

Selenium was discovered by Berzelius in 1817.

Sources. It is found native and also in combination, forming selenides. It is rare. The method of preparation is not important.

Properties. The physical properties of selenium resemble those of sulphur. It shows several allotropic forms. The principal interest attaching to it is that its power of conducting electricity is affected by light. Several forms of apparatus for the electrical transmission of images have been based upon this property.

Gen. Chem. Rel. The compounds of selenium are analogous to those of sulphur; we have

H_2Se Hydrogen selenide.
SeO_2 Selenium dioxide.
H_2SeO_3 Selenous acid.
H_2SeO_4 Selenic   "

The last is capable of dissolving gold.

TELLURIUM, Te, 129.

Tellurium was discovered by Müller in 1782.

Sources. It is found native, and also in union with bismuth, gold, etc. It is rare. The method of preparation is unimportant

Properties. Tellurium has a metallic lustre and pinkish color. It fuses just below a red heat, and at a temperature somewhat higher boils.

Its compounds are analogous to those of sulphur and selenium, and partly to those of the nitrogen group.

Nitrogen Group. This includes boron, nitrogen, phosphorus, arsenic, antimony and bismuth. They are neither strongly positive nor strongly negative, and act generally as triads, but also frequently as pentads, more rarely as monads. Among their compounds are found some of the most active mineral poisons known. Gold and vanadium may also be classed here.

BORON, B, 11.

Sources. Boron occurs only in combination; as boric acid in discharges of steam in the volcanic region of Italy, as sodium (or calcium) borate found in certain lakes in Thibet and upon the surface of the soil in western parts of the United States. Boron was first prepared by Davy in 1807.

Preparation. It is prepared by the action of potassium, sodium or aluminum on boric anhydride, B_2O_3.

Properties. Boron obtained by the action of sodium or potassium is an amorphous, olive-green powder, insoluble in water and combustible. When obtained by the action of aluminum, it is, if pure, in brilliant crystals, closely resem-

bling the diamond and next to it in hardness. A form of
boron analogous to graphite has been supposed to exist, but
it is probably in an impure condition.

Gen. Chem. Rel. Boron is a triad, and may be regarded
as a link between the carbon and nitrogen groups. Its com-
pounds are not strictly analogous to those of any other ele-
ment.

Boric or Boracic Acid, H_3BO_3. This exists, as already
mentioned, in the steam-jets discharged in some volcanic
regions, and some of its salts occur as minerals. It can be
prepared by dissolving borax in warm dilute sulphuric acid
and allowing the solution to cool. Boric acid forms pearly
scales of a bitter taste, soluble in water and alcohol and very
feebly acid. Heated to 248° F. (120° C.), it forms metaboric
acid, HBO_2, and on still further heating it is converted into
boric anhydride, B_2O_3, which fuses to a clear glass. When a
solution of boric acid is boiled, some of the acid passes off in
the steam. It forms salts called borates, many of which are
irregular in composition. Boric acid is an antiseptic.

Tests. Boric acid has a feeble action on litmus, and turns
turmeric paper to a brown-red color. The best test is the
bright green color imparted to flame.

Exp. Shake a little boric acid with alcohol, pour the solution into a
flat dish and ignite. The outer cone of the flame will be tinged with
green.

Compounds of boron with nitrogen, chlorine, bromine and
fluorine are known.

NITROGEN, N, 14.

Sources. Nitrogen constitutes about four-fifths of air, and
occurs in many animal and vegetable tissues, especially in those
performing the higher functions. It also occurs in the min-
eral kingdom in the form of the sodium and potassium nitrates.

The name means " producer of nitre." It was discovered by Rutherford in 1772.

Preparation. The simplest method of preparing nitrogen is to burn out the oxygen from a limited amount of air. A substance of active combustible qualities like phosphorus or sodium is required for this purpose, as common combustibles, like coal, will be extinguished long before the oxygen is exhausted.

Exp. A piece of phosphorus is placed on a small metal basin or block of wood floating upon water, and, being lighted, a bell-jar or wide-mouthed bottle is placed over it. The white fumes of phosphoric anhydride, P_2O_5, soon fill the jar, and are absorbed by the water, which rises to about one-third the height of the jar. The nitrogen so obtained is impure, but shows the properties of the gas. The pure gas may be made by the action of chlorine upon ammonia or by heating a mixture of potassium nitrite and ammonium chloride.

Properties. Nitrogen is a gas without color, taste or smell. It does not, under ordinary conditions, burn or support combustion. It is not poisonous, but will not support life. From this absence of striking properties no interesting experiments can be performed with it. At high temperature and under the influence of electric sparks it will enter in combination with a number of elements. In this way compounds may be formed with oxygen, boron, silicon, carbon, hydrogen, magnesium, etc. It is a little lighter than air, 47 cubic inches weighing 14 grains; 1 litre weighs 1.25 grms. It can be liquefied only by intense cold and pressure.

Gen. Chem. Rel. Nitrogen is generally a pentad; sometimes it acts as a triad, or even as a monad. Although in the free state it shows so little affinity, it forms by indirect means compounds with most of the elements. It is an essential ingredient of all the higher tissues of animals, and exists also in vegetable structures, but not so abundantly. In many compounds it appears to have weak affinity, and these are apt to decompose. Most of the powerful explosives now in

use—gun-cotton and nitro-glycerine, for instance—owe their qualities partly to the nitrogen present.

Tests. Free nitrogen is recognized by its chemical indifference. In combination it is usually detected by being converted into ammonia by the action of alkalies.

Air. The atmosphere is an intimate *mixture* of about four volumes of nitrogen with one volume of oxygen. It surrounds the earth like a shell or casing, and extends upward to a height which has been variously estimated at from 45 to 200 miles. It is known not to be a compound by several tests, among which are—

1st. It turns brown an alkaline solution of pyrogallin, and reddens nitric oxide, characters which belong only to free oxygen.

2d. Air dissolved in water has its nitrogen and oxygen in a proportion different from that in undissolved air. A chemical compound would have the same composition in both cases.

3d. The composition is not constant nor in exact proportions, either by weight or volume, though approaching very closely the formula N_4O.

The fact that this mixture of gases varies so little in different places and under different circumstances is due to an action called diffusion, which is seen in all gases, in nearly all liquids, and would probably be noticed in some solids if high pressures should be used. All gases mingle with one another, so that sooner or later they produce a uniform mixture in spite of the influence of gravity. The rate of mixture depends on the density of the gas. It is expressed mathematically by saying that the rate of diffusion is *inversely* proportional to the *square root* of the densities. Suppose two gases have densities of 1 and 16 respectively; the square root of the numbers will be 1 and 4, and then by inversion we

find that the lighter gas will diffuse with four times the rapidity of the heavier.

Exp. If we place a wide-mouthed bottle filled with coal gas over a similar one filled with air, as shown in the cut, the two gases will after a short time mix perfectly, and on applying a light an explosion will result. The bottles should have pretty wide mouths, or the explosion might be dangerous. The rapidity with which the odor of coal gas penetrates through a room when a leak in a pipe occurs is an instance of diffusion. If the gas were to obey the usual laws of gravity, it would rise to the ceiling, but it really penetrates to all parts almost the same as if it were escaping into an empty space. In the same way, although nitrogen has a density of 14 and oxygen of 16, they do not form separate layers, but are uniformly and permanently mingled.

Ordinary air contains small quantities of other bodies besides nitrogen and oxygen. It always contains water, carbonic acid and ammonia; frequently compounds of nitrogen and oxygen, and also ozone. Besides these we have dust and the products of animal and vegetable decomposition. The study of impurities of air has received much attention of late years, especially in view of the theory that many diseases are due to living organisms or germs which exist in the air. For these investigations the microscope has been necessary.

The approximate composition of air was first demonstrated by Lavoisier in 1777.

The accurate analysis varies somewhat; the following may be taken as a fair average:

Oxygen,	20.61
Nitrogen,	77.95
Carbon dioxide,	.04
Water,	1.40

Traces of ammonia, nitric acid and marsh gas (CH_4), and, in towns, sulphur compounds.

The uniformity of the composition of air is assisted by the

winds and currents which are continually agitating it. Its chemical properties are those of oxygen in a much diminished degree, on account of the dilution with nitrogen. It was formerly taken as a standard for the specific gravity of gases, but hydrogen is now preferred. 100 cubic inches weigh 30.93 grains; 1 litre weighs 1.29 grammes; 13 cubic feet weigh about 1 lb. At the level of the sea the pressure is, ordinarily, about 15 lbs., and will sustain a column of mercury 760 millimetres, or 30 inches, in height. Water in its natural condition always contains some air in solution.

The capacity of air for holding moisture increases rapidly as the temperature rises. The dryness or dampness of the atmosphere is not due to the actual quantity of moisture in it, but to the amount in proportion to what the air can take up. A cubic foot of air at 30° F. can absorb about 2 grains of water; if it contains a grain and three-quarters, it will therefore be nearly saturated and seem damp; if the temperature rise to 80° F., the capacity for moisture will rise to 11 grains, and under these conditions it would seem dry if it held three grains of moisture, because, although the amount of moisture is nearly doubled, its capacity for moisture has increased over five times. The nearness of air to saturation is called the RELATIVE HUMIDITY. Air saturated with water has a relative humidity of 100; if half saturated, the relative humidity is 50, and so on. When the temperature falls the moisture separates to a greater or less extent, and we have fog, rain or dew, and if the temperature gets below the freezing-point, we have snow or frost.

Changes produced in the Atmosphere.—The changes which affect the chemical composition of the air are important. The respiration of animals and the processes of combustion are continually removing oxygen and introducing water, carbonic acid and more or less organic matter. The decay of animal and vegetable substances introduces various gases, especially ammonia and hydrogen sulphide; and the sulphur

of coal and coal gas furnishes sulphurous and sulphuric acid. The dust which is always floating in the air contains a great variety of substances living and dead, and varies with the locality. The continued removal of oxygen is counterbalanced by the respiration of plants, which, under the influence of light, decompose the carbonic acid, retaining the carbon and giving off the oxygen, especially at the under surface of the leaves. In this way the two great divisions of organic nature sustain each other. This fact is well shown in the construction of the ordinary aquarium, in which animal and plant life are maintained for a long time without renewal of the water. The nitrogen of the atmosphere is very little affected. The ammonia and other gases are gradually oxidized or absorbed by the soil and plants and washed out by the rains. The organic matter also oxidizes, and ozone is supposed to be especially active in this respect.

Ventilation.—When animals are compelled to breathe air in a closed space, it becomes, sooner or later, by the removal of oxygen, incapable of supporting their life, and by the introduction of organic matter it becomes an active agent of disease. As all buildings interfere with the free circulation of the air, the problem of ventilation or proper renewal of the air is a very important one. The products of respiration and combustion are usually lighter, because warmer, than the air around, and tend to rise, and the simplest systems of ventilation take advantage of this fact by arrangements which allow the foul air to escape at the top of the room and fresh air to enter at the bottom. Unfortunately, walls and windows are usually cold; they chill this foul air and interfere seriously with its upward movement. Practically, good ventilation cannot be accomplished without some active mechanical assistance, such as a fan-blower or the draft of a chimney.

Ammonia Gas, Amine, NH_3. This substance has been known in some forms from a remote period. The name is derived from the temple of Jupiter Ammon in the Libyan

desert, which was a point at which the materials for making one of its compounds were collected. Ammonia exists in the air in very small quantities. It is given off in the decomposition of organic matter, especially animal remains, and was originally derived from refuse of this kind. It is also produced by the action of hydrogen on nitric acid. The great source at the present time is the water which has been used for washing the common illuminating gas. The distillation of coal produces considerable ammonia, and this is removed by passing the gas through water. The so-called *ammoniacal liquor* is neutralized with an acid, and the resulting compound properly purified. Ordinarily, hydrochloric acid is used, and the reaction is $NH_3 + HCl = NH_4Cl$. For the preparation of ammonia some compound containing it is decomposed by an alkali. The usual method is by treating a mixture of ammonium chloride, NH_4Cl, with lime; the reaction is

$$2NH_4Cl + CaO = 2NH_3 + H_2O + CaCl_2.$$

By passing the gas over dry lime the water is absorbed and the pure NH_3 is collected.

Ammonia is a colorless gas of a pungent odor. It is absorbed by water in large amounts, one pint absorbing 700 pints of gas and increasing fifty per cent. in volume. This solution exhibits most of the properties of the gas, and is much used under the name of aqua ammoniæ or solution of ammonia.

Exp. The affinity of ammonia for water may be easily shown by filling a flask with the gas and inverting it in water. The ammonia will be rapidly absorbed, and the water will rise into the vacuum thus produced.

Ammonia burns, but not easily. It is lighter than air. 1 litre weighs 0.76 grm.; 47 cubic inches weigh 8.5 grains. It contains one volume of N and three volumes of H, condensed to two volumes. At a temperature of —40° F.

(— 40° C.), or under a pressure of 100 lbs. to the square inch, it condenses to a colorless liquid. This liquid of course evaporates rapidly when the pressure is removed, and produces great cold, which fact has been made use of in machines for making ice.

Gen. Chem. Rel. The important property of ammonia is the strongly alkaline and basic power of its solution in water. This solution has chemical characters so much like those of potassa and soda that it is now generally believed that the ammonia and water have formed a compound which is analogous to the alkalies. Long before the composition of the solution was understood it had received the name of *volatile alkali*, to indicate its resemblance to potassa and soda, which were called fixed alkalies. Berzelius was the first to suggest that the compounds produced by ammonia were formed in the same manner as those of potassium and sodium, these elements being represented by the radicle, NH_4. In this way $NH_3 + HCl$ would produce NH_4Cl; $NH_3 + H_2O$ would produce NH_4HO. NH_4 is called ammonium; its atomicity is monad; it combines with one atom of chlorine and replaces the hydrogen of acids.

The following formulæ show the comparison between the salts of potassium and those of ammonium:

KCl	Potassium chloride.
NH_4Cl	Ammonium chloride.
K_2SO_4	Potassium sulphate.
$(NH_4)_2SO_4$	Ammonium sulphate.
KNO_3	Potassium nitrate.
NH_4NO_3	Ammonium nitrate.
KHO	Potassium hydrate.
NH_4HO	Ammonium hydrate.

As the expression NH_4 makes some confusion in formulæ, it is convenient to use the symbol Am; ammonium chloride will thus be AmCl; ammonium sulphate, Am_2SO_4. The am-

monium theory received considerable support when proposed from the following experiment:

If a piece of sodium be put on the surface of some mercury, and a drop of water added cautiously, an alloy called sodium amalgam will be formed. If this be dropped into a strong solution of ammonium chloride, a remarkable increase in bulk of the amalgam will take place. It was at first supposed that a compound had been formed between the mercury and NH_4 by the following reaction:

Sodium amalgam.		Ammonium amalgam.
$HgNa_2 + 2NH_4Cl$	$=$	$2NaCl + (NH_4)_2Hg$.

The property of forming an amalgam with mercury was regarded as proof of the existence of NH_4. Subsequent research has shown that the supposed ammonium amalgam is nothing but mercury inflated with gas-bubbles. The mass rapidly diminishes in bulk, giving off a mixture of NH_3 and H.

Although this experiment is no longer of any weight, the theory is generally accepted, because it greatly simplifies the study of the ammonium compounds.

Ammonium, NH_4, and ammonium oxide, $(NH_4)_2O$, have not been obtained in a form convenient for study.

Ammonium Hydrate, NH_4HO, has already been mentioned as the supposed result of the solution of ammonia gas in water. It is a clear liquid, lighter than water, corrosive and powerfully alkaline and pungent.

Ammonium Carbonate, $(NH_4)_2CO_3$, is made on the large scale by heating ammonium chloride with chalk. The theoretical reaction would be

$$2AmCl + CaCO_3 = CaCl_2 + Am_2CO_3.$$

The Am_2CO_3, however, decomposes into Am_2O, which escapes and leaves $2Am_2CO_3 + CO_2$, ammonium anhydro-carbonate (see page 44), often called sesquicarbonate, or smelling-salt. It is a white body, soluble in water, and smelling strongly of ammonia. By exposure to air it is converted into acid carbonate, $AmHCO_3$.

10 *

Ammonium Nitrate, $AmNO_3$, is made by saturating nitric acid with ammonia. It is a white solid, very soluble in water. Its chief use is for making nitrous oxide.

Ammonium Sulphate, Am_2SO_4, is obtained by boiling gas-liquor and passing the vapors into sulphuric acid. It is used as a fertilizer and in the manufacture of alum.

Ammonium Phosphates. Of these the most important is microcosmic salt, $AmNaHPO_4$. It is used in blowpipe analyses. Various ammonium phosphates exist in manures, and are valuable fertilizers.

Ammonium Chloride, *Sal Ammoniac*, $AmCl$, is prepared as described on page 111. It is a white solid, crystallizing in cubes, and is very soluble in water. It volatilizes without fusing, and is apparently decomposed into HCl and NH_3. It has many uses in the arts, in medicine and in analytical chemistry.

Ammonium Bromide and *Ammonium Iodide* are used in photography and medicine.

Tests. Any ammonium compound may be detected by heating it with potassium or sodium hydrates or slaked lime. Ammonia gas is quickly evolved, and may be recognized by its odor, alkaline reaction and the white cloud of NH_4Cl produced when it comes in contact with vapors of hydrochloric acid. Ammonium compounds give the precipitates with platinum chloride and tartaric acid similar to those yielded by potassium. The most delicate test for ammonium is Nessler's reagent, a solution made by mixing $HgCl_2$, KI and KHO or $NaHO$. The liquid so formed produces, with very minute quantities of ammonia, a yellow color or a yellowish-red precipitate. One part of ammonia in fifty million parts of water can be easily recognized.

Nitrogen Oxides. Five compounds of nitrogen and oxygen have been obtained:

N_2O	Nitrous oxide, laughing gas.
NO	Nitric oxide (often written N_2O_2).
N_2O_3	Nitrous anhydride.
NO_2	Nitrogen peroxide (often written N_2O_4).
N_2O_5	Nitric anhydride.

The system of names given to these compounds is in some confusion, partly because they were formerly assigned a some- what different composition, and partly because some writers double the formulæ. Thus, NO is often written N_2O_2, and called nitrogen dioxide, because of the O_2 present. NO_2 is written N_2O_4, and called nitrogen tetroxide. The names and formulæ given above are in the main preferable.

Nitrogen and oxygen combine to a limited extent under the influence of high temperature, especially of the electric spark, generally producing one of the higher oxides. The common source of the compounds is decomposition of nitrates, espe- cially nitric acid.

Nitric Acid, *Aqua fortis,* HNO_3. This body has been known in the free state for a long time, and two of its salts, KNO_3 and $NaNO_3$, are found in large amounts as minerals; the former in India, the latter in Peru and Chili. The usual method of obtaining the acid is by the action of strong sul- phuric acid upon the nitrates. For commercial purposes sodium nitrate is used, being the cheaper salt. The reaction is

$$2NaNO_3 + H_2SO_4 = Na_2SO_4 + 2HNO_3.$$

The complete action requires a high temperature. For a lab- oratory experiment it is preferable to use less sodium nitrate, and the reaction will be—

$$NaNO_3 + H_2SO_4 = NaHSO_4 + HNO_3.$$

Exp. Mix in a retort equal weights of powdered sodium nitrate and sulphuric acid; connect the retort with a receiver set in cold water, and apply to the mixture a gentle heat. The nitric acid distils over, and will condense in the receiver as a slightly yellow fuming liquid. The

experiment is best performed in a well-ventilated place. No corks or gum tubing should be used in connecting the apparatus.

Nitric acid thus obtained has the composition HNO_3, but it is difficult to prevent it absorbing water, and it is generally yellow from slight decomposition; when quite pure it is colorless. The ordinary acid has the composition $2H_2O + HNO_3$. It is a strongly acid liquid, decomposing in the light, highly corrosive and poisonous, and of active chemical qualities. Its special value in chemistry is its high oxidizing power. One-half the oxygen which it contains is available, and is given up to a great variety of substances. Some variation occurs in the action, but the effect is in most cases represented thus:

$$2HNO_3 \text{ decomposes into } H_2O + 2NO + O_3.$$
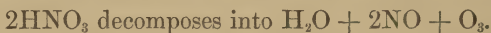
The O_3 is the available oxygen; the NO will be hereafter described as nitric oxide, and its escape is one of the most characteristic evidences of the action of the acid. With some bodies of low affinity the acid acts simply by exchanging its hydrogen for the other substance. Thus:

$$Zn + 2HNO_3 = Zn(NO_3)_2 + H_2.$$
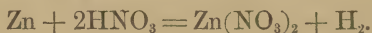
The evolved hydrogen, however, attacks another portion of nitric acid and forms ammonia. This effect is made use of in the detection and estimation of nitrates, especially in drinking-water. The action can be easily observed by making a mixture of zinc and dilute sulphuric acid, and adding nitric acid. The hydrogen, which is copiously evolved when the sulphuric acid alone is present, slowly diminishes upon adding the nitric acid, and may even cease to come off. After a time a perceptible amount of ammonia will be formed. The action is even better produced when we dissolve aluminum foil in a solution of caustic alkali mixed with a small quantity of a nitrate.

With bodies which do not easily displace hydrogen the action of nitric acid is highly characteristic. It exerts the special oxidizing effect referred to above. Two molecules of the acid give up three of oxygen, and form oxides which may or

may not be basic according to the nature of the body or the quantity of the oxygen taken. If basic oxides are formed, they will unite with another portion of the acid to form nitrates. We may suppose, in the case of copper, the following to take place:

$$Cu_3 + 2HNO_3 = 3CuO + H_2O + 2NO.$$

The NO escapes as a gas; the CuO immediately acts upon and neutralizes six additional molecules of nitric acid,

$$3CuO + 6HNO_3 = 3CuN_2O_6 + 3H_2O.$$
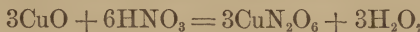
The complete reaction, therefore, is

$$3Cu + 8HNO_3 = 3Cu(NO_3)_2 + 4H_2O + 2NO.$$
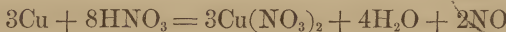
Tin gives the following:

$$Sn_3 + 4HNO_3 = 3SnO_2 + 2H_2O + 4NO.$$

SnO_2, not being basic, does not combine with any nitric acid. Many organic bodies are oxidized in this manner, and are often colored yellow by the action.

Another action of nitric acid is its power of forming substitution compounds. The study of these bodies belongs to organic chemistry, but a single illustrative reaction may be here given. When benzene, C_6H_6, is treated with strong nitric acid, one atom of hydrogen is removed and one molecule of NO_2 put in its place. We have

$$C_6H_6 + HNO_3 = C_6H_5(NO_2) + H_2O,$$
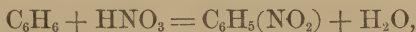
and the body so formed is called *Nitrobenzene.*

A mixture of nitric and sulphuric acids is often used for such effects.

Very strong nitric acid fails to act upon some substances which are readily attacked by the more dilute forms.

Tests. The strong acid may be recognized by its odor and its power of producing yellow stains on organic matter.

Free nitric acid produces the following marked reactions:

Morphia and brucia are colored red; copper is dissolved, with the production of red fumes of NO_2. When the acid is in combination a few drops of sulphuric acid should be added. As all normal nitrates are soluble in water, no precipitation test for the acid is known.

Nitrous Oxide, N_2O, laughing gas, often called nitrogen monoxide. This body was discovered by Priestley in 1776. It is best obtained from ammonium nitrate, which, when carefully heated, decomposes completely into nitrous oxide and steam,

$$NH_4NO_3 = N_2O + 2H_2O.$$

Nitrous oxide is a colorless, odorless gas, with a somewhat sweetish taste. It can be breathed for a short time without injury, and produces a transient intoxication. When the gas is inhaled in quantity, a short insensibility is produced, during which brief operations, such as teeth extractions and opening of abscesses, may be performed. It supports the burning of ordinary combustibles almost as well as oxygen. It is somewhat soluble in water. At a pressure of fifty atmospheres it becomes a colorless liquid, and is now sold in this form compressed in strong metal cylinders.

Nitric Oxide, NO, often called nitrogen dioxide and written N_2O_2. This gas is the usual product of the action of nitric acid as an oxidizer. It is conveniently prepared by the action of nitric acid on copper. The reaction is

$$Cu_3 + 8HNO_3 = 3Cu(NO_3)_2 + 4H_2O + 2NO.$$

Some N_2O is often produced in this experiment.

NO is a colorless gas, very difficult to reduce to the liquid form. Its most marked property is that as soon as it is brought in contact with oxygen, it instantly absorbs one atom, becoming NO_2 and turning brownish-red. This body, NO_2, parts with the second atom of oxygen rather easily, and leaves the NO ready to absorb more oxygen. In consequence of this property nitric oxide is used as a sort of carrier of

oxygen, particularly in the manufacture of sulphuric acid. In the ordinary method of making nitric oxide a portion is always converted into NO_2; this can be removed by collecting the gas over water, which absorbs the NO_2.

Nitrogen Peroxide, or **Nitric Peroxide,** NO_2. This body has various names, owing to uncertainty in its chemical relations. It has been called nitrogen tetroxide (being written N_2O_4 by some chemists) and hyponitric acid, and also by other less common names. The proper name would be nitrogen dioxide, to correspond to the formula NO_2, but as this name has, unfortunately and very unwisely, been given by some writers to nitric oxide, it is probable that the rather unsatisfactory name, nitric peroxide, will have to be retained.

NO_2 is a brownish-red gas, of which the color depends somewhat on the temperature; at about $14°$ F. ($-10°$ C.) it condenses to a liquid. It is readily absorbed by water, which decomposes it, producing various compounds according to the proportions used.

The following experiments illustrate the general relation of the nitrogen oxides:

1. Mix in a glass bottle or retort some copper turnings with moderately strong nitric acid. The evolved gas will have a red color, and consist of a variable mixture of N_2O, NO and NO_2. Pass the gas into a jar of water inverted over the pneumatic trough, and it will be found that a colorless gas is collected. This is because the NO_2 is absorbed by the water.

2. Allow a few bubbles of air to pass up into the jar; a red color is at once produced, which soon disappears, the water rising slightly in the jar. This effect is due to the conversion of the NO into NO_2, and the subsequent absorption of the latter by the water. The experiment may be repeated until all the NO is oxidized, and the residual gas is a mixture of N_2O and N.

3. If pure nitric oxygen and pure oxygen be used in this experiment,

the relation by volume will be easily shown. Two pints of nitric oxide contain one pint of N and one of O. If mixed with an additional pint of oxygen, NO_2 will be formed, and complete condensation will ensue.

Nitrous anhydride, N_2O_3, and nitric anhydride, N_2O_5, are unimportant, as is also nitrous acid, HNO_2; but a few of the nitrites are of importance. Ammonium nitrite occurs in rain-water, and other nitrites are found in river- and well-water. They act both as oxidizing and reducing agents.

PHOSPHORUS, P, 31.

Sources. Phosphorus occurs principally as calcium phosphate, which exists in bones and teeth, in many minerals and in soils. Various phosphates also exist in the fluids of the animal body. It was discovered by Brandt in 1669. The name means " carrier of light."

Preparation. Phosphorus is prepared from bones, which contain from one-third to two-thirds their weight of calcium phosphate. The bones are deprived of their animal matter, and the *bone-ash* thus left is treated with sulphuric acid, by which a soluble acid calcium phosphate is formed and much calcium sulphate is deposited. The liquid is then concentrated, mixed with charcoal and sand and heated. Calcium silicate is produced, the charcoal takes the oxygen, and the phosphorus distils over. The reactions are complicated; the following is to be considered as merely an indication of the change, and not the exact formula:

Bone ash.
$$Ca_3(PO_4)_2 + 2H_2SO_4 = 2CaSO_4 + CaH_4(PO_4)_2.$$

Acid cal. phosph.   Sand.
$$CaH_4(PO_4)_2 + SiO_2 + C_5 = CaSiO_3 + 2H_2O + 5CO + P_2.$$
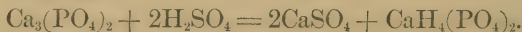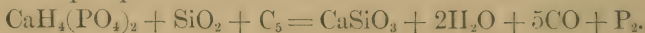
Other methods of manufacture are known, but need not be

described. The distilled phosphorus is collected in water and cast in sticks.

Properties. Phosphorus when freshly prepared is a colorless, almost transparent solid, soft as wax; when kept for some time, especially in the light, it becomes brownish, opaque and harder. It takes fire easily, and is usually kept under water. It burns with a bright flame, producing white clouds of phosphoric anhydride, P_2O_5. Exposed to the air at low temperature, it can still undergo a slow combustion, producing P_2O_3; it is then luminous in the dark. It is insoluble in water, but dissolves in oils and in carbon disulphide. It is extremely poisonous, death having occurred from much less than $\frac{1}{4}$ grain. Phosphorus melts at 111° F. (43° C.), and boils at 550° F. (288° C.), producing a vapor which according to analogy should be 31 times as heavy as hydrogen, but is really 62 times as heavy. By keeping it at a temperature of 450° F. (232° C.) for days in a closed vessel, phosphorus is converted into the

Amorphous or *red phosphorus*, an allotropic form which is red, insoluble in carbon disulphide, difficult to burn, non-poisonous, and shows many other minor differences. Its composition is the same. This change is also produced by adding a small quantity of iodine to common phosphorus.

The uses of the element in matches and as a medicinal substance are well known.

In all experiments with it great care must be taken, as it is easily inflamed and produces one of the most severe forms of burns known. It should be handled with a pair of forceps and cut or divided only under water.

Gen. Chem. Rel. Phosphorus acts as a triad or pentad; its affinities in the free state are very high. It is a powerful reducing agent. Its important properties are shown in experiments given in the description of other elements.

The formation of the amorphous variety may be noticed by adding a few grains of iodine to a small fragment of common phosphorus.

Tests. Phosphorus in the free state is easily recognized. In very minute quantity it is detected by its luminosity when distilled in a dark room.

Hydrogen Phosphide, PH_3, **Phosphine.** This body is formed under conditions analogous to those which produce ammonia; that is, when its elements are brought together in the nascent state (page 37). When a solution of caustic alkali is boiled with phosphorus, water is decomposed and hydrogen phosphide is formed.

Exp. Pieces of phosphorus are put into a small retort, which is then *nearly filled* with a strong solution of caustic soda, and arranged so that the end of the tube can be quickly put below the surface of water in a basin. The retort is heated carefully, and when the mixture boils the hydrogen phosphide will come off. The end of the retort should then be put below the water. The gas passes through the water, and burns on reaching the air. The experiment requires care, and should be performed only by those who have some acquaintance with such manipulations. When the action is over it is best to simply leave the apparatus to itself, as a slight explosion generally occurs.

The reaction is

$$3NaHO + 3H_2O + P_4 = 3NaH_2PO_2 + PH_3.$$
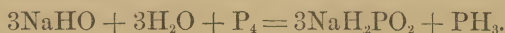
NaH_2PO_2 is sodium hypophosphite. Hydrogen phosphide is a colorless gas of a disagreeable odor. As ordinarily made it is spontaneously inflammable, but this is due to the presence of a small quantity of the vapor of a liquid phosphide, PH_2. If this latter be removed by passing the fresh gas through a tube placed in a freezing apparatus, the power of inflaming spontaneously is lost. A solid phosphide, apparently P_2H, is also known.

The inflammable gas can also be prepared by the action of water on calcium phosphide.

Hydrogen phosphide has no alkaline properties, but forms many compounds analogous to those formed by ammonia.

Compounds of Phosphorus with Oxygen.—Only two compounds are definitely known. These are:

P_2O_3 Phosphorous anhydride.
P_2O_5 Phosphoric anhydride.

Phosphorous Anhydride, P_2O_3, is produced by the slow oxidation of phosphorus. When this takes place in ordinary air, water is absorbed and phosphorous acid, H_3PO_3, is formed. This is a powerful reducing agent, but is not an important body.

Phosphoric Anhydride, P_2O_5. This is the product of the active combustion of phosphorus. It is easily produced by burning phosphorus in the air. It forms white, snow-like fumes, which rapidly absorb water.

Exp. Ignite a small piece of phosphorus on a common plate and cover it with a jar. The white clouds of P_2O_5 soon begin to settle on the plate, and when the combustion is finished throw a few drops of water on the white deposit; a hissing sound will be heard, due to the energetic union of the anhydride with the water.

Phosphoric anhydride is a white, snow-like solid, having a very high affinity for water. It is capable of uniting with water in at least three proportions, forming different bodies; very few anhydrides show this power, and consequently the chemistry of these phosphoric acids is more than ordinarily complicated. When the anhydride is mixed with water as in the above experiment, it combines with one molecule of water and forms HPO_3. If, however, the acid be obtained from any of the phosphates found in nature, it has the formula H_3PO_4; and this can be produced from the anhydride by heating with three molecules of water. An intermediate acid, $H_4P_2O_7$, is also known. The relation between these different forms is shown in the following equations:

$P_2O_5 + H_2O = 2HPO_3$ Metaphosphoric acid.
$P_2O_5 + 2H_2O = H_4P_2O_7$ Pyrophosphoric "
$P_2O_5 + 3H_2O = 2H_3PO_4$ Orthophosphoric "
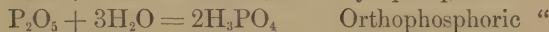
The last acid is the one that yields all the natural phosphates; the syllable "ortho" means "regular," and signifies that this is the regular form of the acid. The syllable

" pyro " means " fire," and signifies that the second acid, $H_4P_2O_7$, is obtained by the action of heat upon some of the natural phosphates. By the further action of heat the metaphosphoric is obtained.

Properties and Gen. Chem. Rel. of the phosphoric acids. Metaphosphoric and pyrophosphoric acids are of comparative unimportance. They are artificial products of the laboratory. The first mentioned is distinguished by a power of coagulating albumen. It is to be particularly noticed that although the three phosphoric acids differ in oxygen, the termination " ic " is not changed. This is because they are all formed from the same anhydride; the difference in oxygen is due to the amount of water. The number of salts formed by each acid is in proportion to the number of molecules of water which it has taken up.

Metaphosphoric Acid, produced by adding *one* molecule of water, gives *one* series of salts:

$NaPO_3$　　　Sodium metaphosphate.
$Ca(PO_3)_2$　　Calcium　　　　"
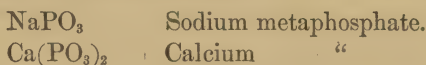
Pyrophosphoric Acid, produced by adding *two* molecules of water, gives *two* series of salts, acid and normal:

$Na_2H_2P_2O_7$　　Acid sodium pyrophosphate.
$Na_4P_2O_7$　　　Sodium　　　　"
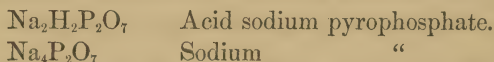
Orthophosphoric Acid, produced by adding *three* molecules of water, gives *three* series of salts, di-acid, acid and normal.

NaH_2PO_4　　　Di-acid sodium orthophosphate.
Na_2HPO_4　　　Acid　　"　　　　"
Na_3PO_4　　　Sodium orthophosphate.
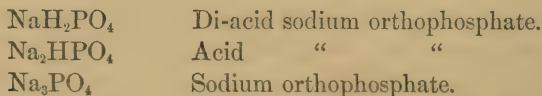
The phosphates of the potassium group are soluble in water. Almost all others are insoluble in water, but soluble in acids.

Tests. The detection of metaphosphoric and pyrophosphoric acids is not often required, but orthophosphoric acid

is a body very often encountered in analysis. Silver nitrate produces with it a yellow precipitate soluble in ammonia. A mixture of magnesium sulphate, ammonia and ammonium chloride gives a white precipitate soluble in acids. A solution of ammonium molybdate in nitric acid gives a bright yellow precipitate insoluble in acid. This is a very delicate test.

The compounds of phosphorus with other elements are of limited importance. With sulphur it gives a variety of combinations. The action between the two elements may be shown by igniting a piece of phosphorus, placed in the midst of a mass of finely-powdered sulphur. A quick combustion, producing a large flame, at once occurs. The phosphorus sulphide first formed is probably quickly consumed.

Two phosphorus chlorides are known, PCl_3 and PCl_5. They have been much used in researches in organic chemistry. Compounds with bromine and iodine are also known.

ARSENIC, As, 75.

Sources. Arsenic occurs in the free state and as sulphide, also in combination, especially with nickel, cobalt and iron. It is rather abundant, and exists in small amounts in many minerals.

Preparation. Arsenic is prepared by deoxidizing arsenous anhydride by charcoal,

$$As_2O_3 + C_3 = As_2 + 3CO.$$
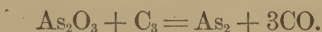
It was first prepared by Brandt in 1733.

Properties. Arsenic, when freshly prepared, is a steel-gray, brittle mass with a decided lustre. It tarnishes somewhat in the air, and passes into vapor at about 356° F. (180° C.) without fusing. The vapor has a density twice as great as analogy would require. Heated in contact with air, it oxidizes to arsenous anhydride, and develops a garlicky odor. It is not dissolved by any simple solvent.

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Gen. Chem. Rel. Arsenic is similar to nitrogen and phosphorus in its chemical relations.

Tests. These are given under arsenous anhydride.

Arsenetted Hydrogen, Arsine, AsH_3. This body is analogous to ammonia, and, like it, is produced when its elements are brought together in the nascent state. Its formation is one of the most delicate tests for arsenic. The usual method of preparation is to liberate hydrogen in a solution of arsenous anhydride. It is a combustible gas of disagreeable odor and excessively poisonous.

A solid arsenous hydride of uncertain composition is known.

Compounds of Arsenic with Oxygen.—These are two :

As_2O_3 Arsenous oxide or anhydride.
As_2O_5 Arsenic   "   "

Arsenous Anhydride, Arsenous Oxide, White Arsenic, As_2O_3. This is the substance generally called arsenic. It is obtained by roasting arsenical ores in a current of air, and presents itself in commerce in two varieties:

(*a*) The vitreous form, transparent and colorless at first, but afterward becoming yellowish and porcelain-like ;

(*b*) A pulverulent form, which is distinctly crystalline.

Properties. Arsenous anhydride is a white solid, odorless and tasteless, and dissolving with difficulty and only in small amounts in cold water ; the solution is feebly acid, and is supposed to contain arsenous acid, H_3AsO_3. Hot water is a more active solvent, but the amount dissolved is dependent on many conditions and variously stated by different observers. As a rough statement it may be said that a fluidounce of cold water will dissolve about one grain, and the same amount of water if kept for one hour at the boiling-point will take up about forty grains. In acid and alkaline solutions it dissolves much more readily. Heated to 380° F. (193° C.), the solid passes into vapor without fusing, and if allowed to condense

produces brilliant, transparent crystals. It is intensely poisonous in all its forms, two or three grains being a fatal dose. Arsenous anhydride is used in making opaque white glass, in various solutions for preserving animal skins, and in the manufacture of colors. Its frequent occurrence and poisonous qualities have made its properties and tests of great importance. The following is a brief summary of the methods used.

1. *Reduction Test.*—This depends on the conversion of the arsenical compounds into the elementary arsenic. A small quantity of powdered white arsenic is mixed with some dried potassium ferrocyanide and heated in a narrow glass tube. The elementary arsenic is set free, rises in vapor, and condenses on a cooler portion of the tube, as a dark steel-gray but rather lustrous layer called the *arsenical mirror*. If this deposit be heated, it may be driven farther along the tube, and will finally oxidize and produce the garlicky odor.
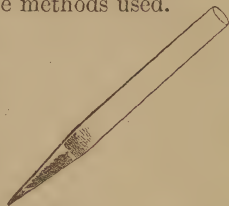
2. *Sublimation Test.*—Arsenous oxide heated alone passes quickly into vapor, and by allowing this vapor to condense upon a *slightly warmed* part of the tube fine crystals are formed. Under the microscope these crystals are seen to be octahedral; that is, consist of eight triangular faces, though they are rarely completely formed. Very minute quantities of arsenic can be recognized by this test.
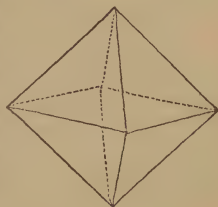
3. *Reinsch's Test.*—This is the most valuable test, because it can be applied to impure mixtures, as the contents of a stomach. A small quantity of water is put into a wide test-tube or porcelain basin; some hydrochloric acid is added; a piece of clean copper is put in and the water brought to boiling. A few drops of a solution of arsenic are now added, and in a few seconds a rather dull steel-colored deposit of copper arsenide forms on the copper. When this deposit has become rather dense, the copper is taken out, dried with filter-paper, rolled up into small bulk and placed in the end of a small glass tube. Heat being applied, the arsenical deposit is oxidized and volatilized, forming octahedral crystals of arsenous anhydride.

4. *Marsh's Test.*—This depends on the power of nascent hydrogen to form AsH_3. The hydrogen is obtained either by the action of sulphuric acid upon zinc or magnesium, of sodium amalgam on water, or by a

current of electricity. The test is very delicate, but requires great skill in its manipulation. The simplest method of performing it consists in mixing zinc and dilute sulphuric acid in a gas bottle, allowing the hydrogen to flow for some time, and then introducing a small quantity of arsenical solution. The arsine, AsH_3, begins at once to come off; the flame of the hydrogen becomes livid and gives off fumes of arsenic. If a cold porcelain plate be held in the flame, an arsenical soot will be deposited as a brown shining stain. If the tube which is conducting the current be heated, the gas will be decomposed and a similar stain formed within the tube. The stains may be identified as arsenic by the fact that they are—

(a) easily volatile;

(b) soluble in a solution of bleaching-powder;

(c) capable of producing octahedral crystals of As_2O_3.

Three tests, known as the liquid tests, are applicable only to pure solutions of arsenous anhydride. They are—

1. Hydrogen sulphide produces a lemon-yellow precipitate of arsenous sulphide, As_2S_3.

$$As_2O_3 + 3H_2S = As_2S_3 + 3H_2O.$$

A few drops of hydrochloric acid facilitate the action.

2. Silver nitrate, made alkaline by ammonia, gives a yellow precipitate of silver arsenite.

3. Copper sulphate, made alkaline by ammonia, gives a green precipitate of copper arsenite.

Arsenic Anhydride, As_2O_5. This is produced by oxidizing arsenous anhydride with nitric acid. It forms, with water, arsenic acid, H_3AsO_4, which is used as an oxidizing agent in the manufacture of aniline colors. This use has been supposed to account for the cases of skin irritation which have been occasionally observed to follow the wearing of fabrics dyed with these colors, but it is very doubtful if any arsenic remains in the manufactured article.

Arsenic acid forms salts called arsenates. It is usually tested by first converting it into arsenous acid. Three forms of arsenic acid are known, corresponding to the three forms of phosphoric acid.

Compounds of Arsenic and Sulphur.—Three of these are known:

As_2S_2	Arsenous disulphide, realgar.
As_2S_3	Arsenous sulphide, orpiment.
As_2S_5	Arsenic "

Realgar is a brick-red solid, easily volatile. It is found as a mineral, and may also be produced artificially. It has little practical importance.

Orpiment, King's yellow, is found as a mineral, and is easily produced artificially by the action of hydrogen sulphide upon arsenous anhydride.

$$As_2O_3 + 3H_2S = As_2S_3 + 3H_2O.$$

It is a bright yellow solid, fusible and volatile, soluble in alkalies, but insoluble in water and dilute acids. It is often obtained in the process of testing for arsenic, and in the arts is used as a pigment.

Arsenic sulphide is unimportant.

Arsenic forms chlorides, bromides and iodides, but they need not be described.

ANTIMONY, Sb, 122.

Sources. Antimony occurs sometimes in the free state, but generally as sulphide, Sb_2S_3. It was discovered by Basil Valentine in the fifteenth century. It is also called *Stibium.*

Preparation. By melting the roasted antimony sulphide with charcoal and sodium carbonate. The reaction is analogous to that occurring in the preparation of arsenic.

Properties. Antimony is bluish-white, brittle, generally highly crystalline and of brilliant lustre. It fuses at 842° F. (450° C.), and volatilizes at a red heat. On cooling from

the melted condition it expands somewhat, and some of its alloys retain this property, for which reason it is used in type-metal and other alloys which must take sharp casts. Like arsenic, it is not soluble in any simple solvent.
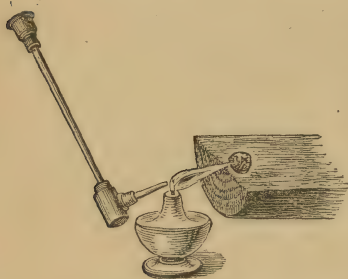
Exp. Heat a piece of antimony about the size of a large shot on a piece of charcoal by means of the blowpipe, and when the mass is in perfect fusion drop it from a height of several feet upon a large sheet of paper. The globule will splash and throw little globules in all directions, each of which leaves a train of antimony oxide.

Gen. Chem. Rel. The chemical relations of antimony are much like those of arsenic, phosphorus and nitrogen, but, unlike them, it forms an oxide which is slightly basic.

Tests. Antimony is detected by tests similar to those of arsenic. The distinctive differences are:

1. The sublimate of free antimony cannot be obtained by the reduction test unless a very high temperature be used.

2. The antimony oxide cannot be volatilized except by a high heat, and does not usually form octahedral crystals, but these have been obtained under certain conditions.

3. The copper slip in Reinsch's test becomes covered with a bluish or violet deposit, which gives a sublimate only with great difficulty.

4. In Marsh's test a much darker spot is obtained; it is volatilized with difficulty, and not dissolved by a solution of bleaching-powder.

5. The liquid tests give no result except with hydrogen sulphide, which produces an *orange-red* precipitate.

Stibine, Antimonetted Hydrogen, SbH_3, resembles the

corresponding arsenic compound and is produced under similar conditions. It has not been obtained pure. The distinction between it and arsine is given above.

Compounds of Antimony with Oxygen.—These are—

Sb_2O_3 Antimonous oxide, or anhydride.
Sb_2O_5 Antimonic " "

An intermediate oxide, Sb_2O_4, is known, but it is generally regarded as a compound of the other two.

Antimonous Oxide, Sb_2O_3. This is found as a mineral, and is also readily prepared by burning antimony in the air. It is like As_2O_3 in many of its chemical relations, but is insoluble in water, less volatile, and shows some power of combining with acids to form salts. When boiled with a solution of cream of tartar (acid potassium tartrate) antimonous oxide loses one atom of oxygen, and dissolves, forming *tartar emetic*, potassium antimony tartrate. This compound is the most familiar preparation of antimony, as, unlike most of the compounds, it dissolves in water without decomposition. The composition is exceptional; acid potassium tartrate is $KHC_4H_4O_6$, and the reaction with antimonous oxide is

$$2KHC_4H_4O_6 + Sb_2O_3 = 2K(SbO)C_4H_4O_6 + H_2O.$$
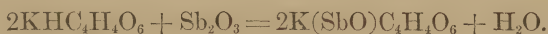
The formula is generally as if the SbO replaced the hydrogen as a monad radicle, but if written graphically it is more satisfactory:

$$K-(C_4H_4O_6)-Sb=O.$$

$C_4H_4O_6$, the radicle of tartaric acid, is a dyad. Boron and arsenic may take the place of antimony in this compound.

Antimonic Oxide, Sb_2O_5, forms two acids corresponding to the meta- and pyrophosphoric acids. Unfortunately, the names of the antimony acids have been misplaced; $HSbO_3$ has been called antimonic acid, and $H_4Sb_2O_7$ metantimonic. This latter name should be given to $HSbO_3$, and $H_4Sb_2O_7$

should be called pyrantimonic acid. Pyrantimonic acid is remarkable for forming the only sodium compound insoluble in water. The orthantimonic acid, H_3SbO_4, has not yet been obtained.

Antimony forms compounds with chlorine, bromine and iodine analogous to those of phosphorus and arsenic. They are mostly decomposed when mixed with large quantities of water, yielding at first an impure, finally a pure, oxide. With antimonous chloride we have

$$3SbCl_3 + 3H_2O = SbCl_3Sb_2O_3 + 6HCl.$$

The oxychloride, $SbCl_3Sb_2O_3$, becomes finally converted into pure antimonous oxide.

Antimony Sulphides. Two are known:

$$Sb_2S_3 \qquad \text{Antimonous sulphide.}$$
$$Sb_2S_5 \qquad \text{Antimonic sulphide.}$$

Antimonous Sulphide is the principal ore of antimony. It is found as a shining, gray, crystalline mass, fusible and easily oxidized by heating in the air. Hydrochloric acid dissolves it easily, forming antimonous chloride and hydrogen sulphide.

$$Sb_2S_3 + 6HCl = 2SbCl_3 + 3H_2S.$$

On the other hand, a current of hydrogen sulphide passed into antimony solutions produces the antimonous sulphide as an *orange-red* precipitate, which by heating becomes like the natural form.

Antimonic Sulphide is an orange-yellow body.

The chemical relations of antimony are well shown in its sulphides. Both of them act as anhydrides, and form a series of salts.

$$KSbS_2 \qquad \text{Potassium sulphantimonite}$$

is strictly comparable to

$$KNO_2 \qquad \text{Potassium nitrite.}$$

Antimonic sulphide, unlike the corresponding oxide, forms its salts upon the pattern of the orthophosphates.

$$Na_3SbS_4 \qquad \text{Sodium sulphantimonate}$$

is analogous to

$$Na_3PO_4 \qquad \text{Sodium orthophosphate.}$$

Sodium sulphantimonate is used in photography under the name of *Schlippe's salt.*

BISMUTH, Bi, 210.

Sources. Bismuth is commonly found native; also as oxide and sulphide. It was known to the earlier chemists.

Preparation and Properties. It is easily extracted from its ores by fusion in iron cylinders. It is hard, brittle, reddish-white and distinctly crystalline. It fuses at 507° F. (264° C.), expanding when it solidifies. Its chemical relations are somewhat like those of antimony, but it forms well-marked salts. It is not much affected by the air. Nitric acid dissolves it easily. The bismuth oxides are $Bi_2O_2, Bi_2O_3, Bi_2O_4, Bi_2O_5$.

Bismuth Sesquioxide, Bi_2O_3, the only important oxide, is obtained as a yellowish powder by burning bismuth in the air or by heating the carbonate or nitrate. It acts as a base.

Bismuth Nitrate, $Bi(NO_3)_3$, made by dissolving bismuth in nitric acid, is a soluble, white, crystalline mass. When added to a large volume of water, a white precipitate of bismuth oxynitrate of irregular composition, but generally $Bi(NO_3)_3 + Bi_2O_3$, is thrown down. This powder, ordinarily called *bismuth subnitrate*, is used in medicine and sometimes as a cosmetic. When it is boiled with caustic soda and a solution of glucose, a heavy black powder of free bismuth is formed. This is Boettger's test for sugar.

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Bismuth Chloride, $BiCl_3$, is decomposed by water in a manner similar to the nitrate.

GOLD, Au, 196.7.

Sources. Gold occurs in the free state, often in veins in quartz, often in small grains in sand and gravel; sometimes alloyed with silver, copper or other bodies. Gold was known to the ancients.

Preparation. It is often extracted by washing, which carries off the lighter sand and dirt and leaves the gold. Quartz rock is first ground. It is sometimes extracted by amalgamation, as described under silver.

Properties. Pure gold is bright yellow, soft and heavy (specific gravity, 19.4), capable of being worked into thin plates or wire, and an excellent conductor of heat and electricity. It melts at 1900° F. (1036° C.). It is unaffected by air, water or sulphur, or by ordinary acids even at high temperatures. Its compounds are reduced by heat alone, and by reducing agents in the cold. Chlorine or a mixture of nitric and hydrochloric acid (which contains free chlorine) dissolves it, forming chloride. In the pure condition it is very soft and can be welded in the cold by pressure. Gold-foil is prepared in this form for dentists' use. For jewelry, coin and other articles subjected to wear it is always alloyed with copper or silver. The proportion of alloy is indicated by *carats*, pure gold being 24 carats, 18-carat gold being 18 parts gold and 6 parts alloy. Copper makes a red gold—silver, a green gold. Two sets of compounds are known, aurous and auric, in which the metal is respectively a monad and a triad. The oxides are not bases; one appears to be an anhydride.

$$Au_2O \quad \text{Aurous oxide.}$$
$$Au_2O_3 \quad \text{Auric} \quad \text{``}$$

AuCl   Aurous chloride.
AuCl$_3$ Auric   "

Auric Chloride is produced when gold is dissolved in nitro-muriatic acid. By adding to the liquid, free from excess of acid, some ferrous sulphate, the gold is thrown down as a brown powder, looking like mud. A mixture of stannous and stannic chlorides produces with gold chloride a purple precipitate called *purple of Cassius*, which is used for coloring glass and porcelain.

Thallium, Tl, 204, was discovered by Crookes in 1861. It exists in some varieties of iron pyrites and in some mineral waters. It acts as a monad and triad, and resembles lead in many points; but its compounds are somewhat like those of potassium, and somewhat like those of silver. Sp. gr. 11.8. It forms two oxides, Tl$_2$O and Tl$_2$O$_3$, the former being very soluble in water; two chlorides, TlCl, insoluble in water, and TlCl$_3$, soluble in water; and other salts of similar relations. Thallium compounds give a pure green color to flame.

Vanadium, V 51.3, discovered by Del Reo in 1801, is a rare body, found chiefly in combination with iron and lead. It forms four oxides, VO, V$_2$O$_3$, VO$_2$, V$_2$O$_5$, analogous to those of nitrogen.

Vanadic Anhydride, V$_2$O$_5$, forms salts called vanadates. Lead vanadate is found as a mineral. It yields compounds analogous to metaphosphoric acid, and also forms salts with some of the strong acids. Vanadium has acquired some importance from the possibility of making from it a good indelible ink, but the rarity of its compounds has interfered with this use.

Carbon Group. This includes carbon, silicon, tin, titanium, and provisionally tungsten, zirconium, platinum,

palladium and ruthenium. They are tetrads, neither strongly positive nor strongly negative in character. With the exception of carbon and silicon they form feebly basic oxides. All of them form acid oxides (anhydrides). The group will doubtless be divided when the less common members of it are more thoroughly studied.

CARBON, C, 12.

Sources. Carbon occurs very abundantly in nature, both in the free state and in combination. It is so constant a component of organic bodies that organic chemistry has been rather fancifully called the chemistry of the compounds of carbon. In the tissues of animals and plants it exists in union with hydrogen, oxygen and nitrogen. The various forms of coal and graphite, and certain carbonates, especially of calcium and magnesium, are abundant minerals. Carbon is remarkable for presenting itself under modifications so different that we should hardly suppose them to be of the same composition. These are—

Amorphous Carbon, of which lampblack and charcoal are examples.

Graphite, or *Plumbago*, which is not perfectly pure, and is probably a very old form of coal.

Diamond, which is often chemically pure, but is also found in inferior conditions.

Some properties are common to all these forms. They are all insoluble in all liquids, and infusible and unacted upon by acids and alkalies or by the air at ordinary temperatures. Heated strongly in air or oxygen, they burn, producing CO or CO_2.

A number of impure forms of carbon are also known. The special properties and origin of the forms, pure and impure, need only brief description.

LAMPBLACK is the deposit from smoky flames. It is a soft black substance, used for printing-inks and colors. It generally contains hydrogen.

WOOD CHARCOAL is obtained by heating wood out of contact of air. It contains hydrogen and the mineral substances of the wood.

ANIMAL CHARCOAL is obtained by charring blood, bones and other animal tissues. It is a coarse black powder.

Wood and animal charcoals have great powers of absorption—the former for gases, the latter for organic matters, especially colors and bitter principles.

Exp. If a piece of wood charcoal be weighted so as to sink in some water in a tall jar, and the jar then be placed under the receiver of an air-pump, and the air exhausted, large quantities of gas, principally nitrogen, oxygen and carbon dioxide, will escape from the charcoal. This property of wood charcoal explains its use as a deodorizer. Gases containing hydrogen, sulphur or phosphorus are generally entirely burned up or decomposed when absorbed in this way.

Exp. If a solution of some organic color, such as litmus or cochineal, be filtered through animal charcoal, the color will be partly or wholly removed. Bitter principles, such as strychnia or the bitter of hops, will also be removed. Animal charcoal is extensively used for the decolorization of syrups and vegetable infusions generally.

GRAPHITE, called also plumbago and black lead, is destitute of any absorbent properties, and is used for lead-pencils and for crucibles.

DIAMOND is a crystalline form of carbon, and has been produced artificially, though in very minute form. Its origin is unknown, and it is generally found in soil which has been transported by water. It is the hardest substance known, and in the impure and discolored forms it has been used with great advantage for the drilling and cutting of stone.

The secondary properties, such as specific gravity, color and hardness, are different in the various forms of carbon.

12 *

The impure forms of carbon are the varieties of coal.

Coal has been formed from vegetable matter by a slow process of decay, mostly under water, by which the hydrogen and oxygen are in great part removed, and the carbon by pressure made compact. Bituminous or soft coals are thus produced. They always contain hydrogen and oxygen, and when heated evolve a variety of gases which have high illuminating power, and constitute ordinary coal gas. Remains and impressions of plants are found in such coal. Coke is the residue after heating the coal. In the north-western part of the United States coal of comparatively recent origin occurs. Anthracite coal is much harder, and has very little hydrogen. It yields no gas on heating. It is practically a compact coke. The anthracite coalfields of Pennsylvania are the most valuable in the world, very few deposits like them being known.

Gen. Chem. Rel. Carbon is a tetrad, and combines with many elements.

Tests. Carbon in the free condition is easily recognized by its infusibility and combustibility, and by producing carbonic acid. Most of its complicated compounds yield a black residue of charcoal when strongly heated or mixed with strong sulphuric acid.

Compounds of Carbon with Hydrogen.—Hydrogen and carbon combine in many proportions, forming a great variety of bodies. Some of the substances are produced by elaborate chemical operations, some by the natural processes of decay, many important ones by the action of heat on organic substances, especially coal and wood.

COAL GAS. When bituminous coal is heated in a closed vessel, a large amount of gas (about five cubic feet to the pound in good samples) is given off. This gas is contaminated with tar, ammonia and sulphur compounds. The tar deposits on cooling, the ammonia is removed by water, and the sulphur by lime; and thus purified it constitutes illuminating

gas, which is a variable mixture of hydrogen, marsh gas, CH_4, olefiant gas, C_2H_4, and other gases.

Compounds of Carbon with Oxygen.—The important ones are—

Carbon monoxide	CO.
Carbon dioxide $\big\}$	CO_2.
Carbonic anhydride	
Oxalic anhydride	C_2O_3.

The last is not known in the free state.

Carbon Monoxide, Carbonic Oxide, CO. This is produced when carbon is burned in a deficient supply of air, as in stoves with defective draft and in the large furnaces for reducing and working iron, in which an excess of fuel is purposely maintained. When steam is thrown upon burning coal, the reaction $H_2O + C = CO + H_2$ occurs, and the resulting mixture is available as a gaseous fuel, or may be impregnated with vapors of benzine or gasoline and used as a source of light. In making carbon monoxide for experimental process these methods are unsuitable, and application is made of the fact that the action of sulphuric acid upon oxalic acid or upon potassium ferrocyanide gives rise to the gas.

Exp. Mix in a rather large flask some crystallized oxalic acid with its own bulk of sulphuric acid, and heat carefully. The mass will soon begin to foam and give off a mixture of CO and CO_2. By passing this through lime-water or caustic soda, pure carbon monoxide may be obtained. The reaction is

$$H_2C_2O_4 + H_2SO_4 = (H_2O + H_2SO_4) + CO_2 + CO.$$
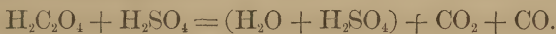
The sulphuric acid and water form a sort of combination.

If potassium ferrocyanide be used in the proportion of about 1 part by weight to 13 parts of sulphuric acid and 1 of water, a large volume of the pure gas may be obtained. The reaction is a complicated one, and the mixture swells considerably, so that a large flask must be used.

Properties. Carbon monoxide is a colorless, odorless, tasteless gas, of decidedly narcotic-poisonous properties. It

is a little lighter than air. It burns easily with a clear blue flame. It is an unsaturated molecule, and will combine with chlorine and some other elements.

Carbon Dioxide, Carbonic Anhydride, CO_2, generally occurs in union with H_2O, forming H_2CO_3, carbonic acid, an abundant substance occurring in air and water. Some of its salts, especially calcium and magnesium carbonates, are common minerals.

Carbonic Acid is produced in a great variety of ways:

1. By the respiration of animals;

2. By ordinary combustion;

3. By fermentation and decay;

4. By decomposition of carbonate, either by heat or by acids. The last method is made available for its production on the small scale.

Exp. Put some fragments of marble or chalk ($CaCO_3$) into a retort or gas bottle, and add some moderately strong hydrochloric acid. The escape of gas occurs at once, and it may be collected by downward displacement or over water. Sulphuric acid does not answer so well, as it soon forms an insoluble mass in the bottle.

The reaction is

$$CaCO_2 + 2HCl = CaCl_2 + H_2O + CO_2.$$

It is unimportant whether we regard the water and CO_2 as separate or united. By passing the escaping gas over dry calcium chloride or strong sulphuric acid the pure CO_2 may be collected. Its properties are substantially those of H_2CO_3.

Properties. Carbonic acid is a colorless gas of a somewhat sharp taste. It is soluble at ordinary pressure in its own bulk of water, and the solubility is increased in regular proportion to the pressure. It is about fifty per cent. heavier than air, and may therefore be easily collected by running the delivery-tube to the bottom of the jar. 1 litre weighs 2.07 grms.; 47 cubic inches weigh 22 grs. It can be liquefied by a pressure of 800 lbs. to the inch, and freezes at

—70° F. (—56° C.).   It does not support animal life or ordinary combustion; but bodies of high affinity, if already in active combustion, will decompose it and continue to burn. In this way red-hot coal will produce the following reaction : $C + CO_2 = 2CO$, which accounts for the production of carbon monoxide in ordinary stoves.

Exp. A lighted taper put into the gas is instantly extinguished, but a slip of ignited magnesium will continue to burn and deposit carbon. The reaction is $Mg_2 + CO_2 = 2MgO + C$.

Exp. Collect the gas in a tube over water, introduce a small piece of caustic soda, and quickly cork the tube.   After a few moments' shaking, open the tube with the mouth under water, when the rise of the water will indicate the absorption of the gas.

Exp. Lime-water is instantly rendered turbid by the gas, from the formation of insoluble calcium carbonate, but if the quantity of the carbonic acid is in excess, the precipitate will be redissolved, producing what is ordinarily known as a hard or limestone water.   This reaction is more fully described in connection with the calcium salts.   If the clear solution produced in this experiment be boiled, the excess of CO_2 will be expelled and the precipitate will be reproduced.

Exp. Soap-bubbles blown with the gas sink rapidly in the air, and if blown with ordinary air will float in a jar of the gas.

Exp. A large light vessel, being counterpoised on a delicate balance, will be thrown decidedly out of balance by substituting carbonic acid for the contained air.   By inverting the vessel the gas will escape and the balance be restored.

Exp. Carbonic acid may be easily poured from one vessel to another, and if a lighted taper be put at the bottom of a tall beaker, it will be extinguished by the gas falling over it.

Exp. Place a piece of potassium in a test-tube having the bottom slightly broken out and arranged in connection with a bottle evolving carbon dioxide, as shown in the cut.   Allow the gas to pass for a few moments, that the tube may become filled, and then heat the potassium strongly.   It will take fire, burning with a purple light

and depositing carbon on the side of the tube.
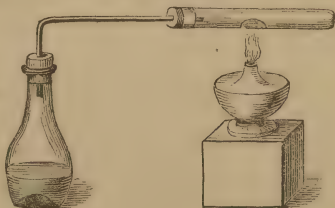
Carbonic acid has, of course, a low diffusive power, and hence a tendency to accumulate at low levels if produced in large amounts. It is found in undue proportions at the bottom of mine-shafts and in fermenting-vats, and cases of suffocation often occur in these places. The usual method of determining whether such places are safe to enter is by lowering a lighted candle; if this continues to burn vigorously, the air is probably safe; if it burns feebly or is extinguished, the air is too rich in carbonic acid.

Gen. Chem. Rel. Carbonic acid may be considered a feeble chemical agent when compared to such bodies as nitric and sulphuric acids, but its continual presence in air and in water makes it one of the important agents in the slow changes which occur in nature. Assisted by the action of frost, it breaks down and renders soluble many kinds of rocks and converts them into soils. Its power of increasing the solvent action of water makes that liquid a very active agent in geological changes. The chemical activity of its solution in water is increased by pressure, and water which is charged with it at great depths often emerges at the surface holding much mineral matter. The release of pressure causes the gas to escape and gives an effervescing water. The mineral matter is at the same time deposited. The ordinary effervescing soda-water is an artificial solution of the gas under pressure. Fermenting liquids owe their effervescence to the same cause, the retention of gas under pressure and its escape when the pressure is released. Its relation to plant-life is very important. Under the influence of light plants decompose it, the carbon being absorbed and the oxygen given off.

Carbonic acid forms a series of salts called the carbonates, most of which are insoluble in pure water. Monads form two salts. Potassium gives us, for instance,

$$KHCO_3 \qquad \text{Acid potassium carbonate.}$$
$$K_2CO_3 \qquad \text{Potassium carbonate.}$$

Dyads give but one salt, as

$$CaCO_3 \qquad \text{Calcium carbonate.}$$

The carbonates are decomposed by almost all acids, and generally by heat.

Tests. Carbonic acid is easily recognized by its rendering turbid a solution of calcium hydrate (lime-water) or barium hydrate (baryta-water). It turns litmus to a wine-red, the blue color being restored on boiling; solution of cochineal is not affected.

Combustion and the Structure of Flame.—The elements carbon, hydrogen, oxygen and nitrogen, and their compounds, form by far the greater part of all the material objects around us, and are the especial elements of the tissues of animals and plants from which our fuel and illuminating agents are directly or indirectly derived. The process of burning is much the same as the process of decay. It is the absorption of oxygen and the formation of carbonic acid, water and free nitrogen, or sometimes of ammonia and nitric acid. The phenomenon of flame attracted the attention of the most ancient investigators, and it was by them considered a form of matter. We know now that ordinary flame is a process; it is gas of some kind in the act of uniting with the oxygen of the air, the operation being attended with the production of light, heat and other forms of force.

Formerly the terms " combustible " and " supporter of combustion " were much used ; carbon, phosphorus and hydrogen being called combustible elements, oxygen and chlorine supporters of combustion. This distinction is now abandoned ; the action is a mutual one, and the supporter of combustion may easily be made the combustible. True flames may be produced in which none of the ordinary agents are used.

The flame of a candle or other body burning free in the air is generally pointed or conical, due to drafts of air which strike the side of the flame and rise, drawing in toward the
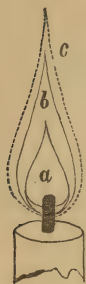
centre. When, as in our gas-burners, the burning body is supplied under pressure, the form of the flame is different, but the different parts are still conveniently called *cones.*

If we examine common gas or candle flame, we find that it consists of three parts:

(*a*) An inner space of a blue color;

(*b*) A shell of brightly luminous particles;

(*c*) A fringe of feebly luminous particles.

The inner cone is the point at which the gas that is burning is produced or escapes. In the flame of a candle or of coal oil the gas is the result of a destructive distillation of the fat or oil. This gas contains carbon and hydrogen. At its outer edge it meets the air; most of the hydrogen is converted into water, the carbon is set free in solid but finely divided condition in union with some hydrogen, and this solid is intensely heated by the combustion of the hydrogen. This is the source of the light, and forms the second cone. The finely-divided matter passes outward and gradually burns, producing the feeble fringe of light, which is the third cone. It is obvious that with bodies which are deficient in carbon, or which are burned in a supply of oxygen sufficient to consume the carbon before it can be set free, very little light will be produced; on the other hand, if the quantity of carbon is large, the flame will not be able to heat it above a red heat, and the supply of oxygen may not be sufficient to burn it up; and we then have a lurid, smoky flame.

The following experiments show the general nature of flame and the manner of studying it:

1. The flame of pure hydrogen gives out very little light. If solid particles of charcoal dust, platinum wire, etc. be introduced, they will become luminous.

2. If a small glass tube be introduced into the centre cone, a quantity of the unburnt gas will pass out, and may be burned at the other end of the tube. The temperature of this interior cone is very low.

3. Alcohol, which contains very little carbon, burns without much light; turpentine, which contains much carbon, burns with a red flame and smoke. By making a mixture of the two a pretty good flame may be obtained.

4. A piece of paper or card, being held for a moment low down in a spirit-lamp flame, will be charred in a ring, showing that the interior of the flame is not burning.

5. Anything which cools the carbon down below its burning-point will cause it to deposit in the solid form; hence the formation of soot or lampblack when flames come in contact with cold surfaces.

6. If a chimney is placed over a smoky flame, the increased draft causes a more abundant supply of air, and the carbon is completely burned. This is the reason for the use of chimneys in coal-oil lamps.

The accompanying illustrations show methods of arranging lamp-chimneys for showing the movements of heated air. The direction of the current may be shown by igniting a piece of brown paper which has been soaked in a strong solution of nitre and dried.
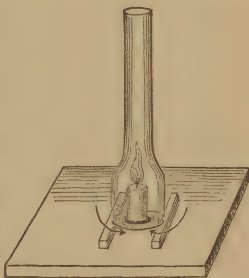
When the flame is disposed in the form of a ring, air being admitted without and within the ring, the arrangement is called an argand burner. The German student's lamp is an excellent form of argand.

7. If a flame be suddenly cooled, as by the introduction of a coil of wire or a sheet of wire gauze, the combustion will cease and the mixture of gas and air will escape. This can be easily shown by putting a piece of wire gauze across a gas flame, when it will be found that the flame will stop at the gauze, but a combustible mixture of gas and air will pass through it. Similarly, the gas may be lighted above the gauze and the flame will not run back. If the gauze becomes hot, the flame will pass through. This principle is made use of in the Davy's safety-lamp for preventing explosions in mines. It consists of a lamp arranged so that no air or gas can get in except through fine gauze. If an explosive mixture finds its way to the

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flame, its combustion is limited to the interior of the lamp, at least for a time.

8. If common coal gas be mixed with air, it will burn with a non-luminous, smokeless flame; and such lamps are now used very largely. In the simplest form, the Bunsen burner, the air is drawn in through openings at the bottom. A great variety of these lamps is now made, and Mr. Fletcher of England has brought the use of gaseous fuel to high perfection, so that it bids fair to come into almost exclusive use.

When a current of air is driven into a flame its temperature is increased. This is the cause of the efficacy of the mouth blowpipe and of blast-lamps.

When mixtures of gas and air are ignited, combustion may occur through the entire mass at once. This constitutes an explosion. Recent research has shown that violent explosions may occur from the rapid ignition of fine particles of coal or flour diffused through the air.

FLAME TESTS.—Many elements give characteristic colors to flames. Such tests are very delicate, and when applied to pure substances very satisfactory. When several colors are present, one color may conceal the other, and thus the test be incomplete. By passing the light through a prism the colors are separated, and each may be recognized. The apparatus for this purpose is called a spectroscope. Observations with it show that most elements give out light which is made up of several colors.

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Carbon Disulphide, CS_2. This body is precisely analogous to carbon dioxide. It is produced by passing vapor of sulphur over red-hot charcoal. It is a colorless liquid, which, when quite pure and in large quantity, has a rather pleasant odor, but when impure, and especially when diffused through the air in small quantity, is quite disagreeable. It has high dispersive power—that is, separates widely the various colors of the spectrum—and is often used in optical apparatus. It is very volatile and inflammable, and has high solvent pow-

ers, dissolving sulphur, phosphorus and most oils and fats, and is much used for such purposes. Its vapor will take fire much below a red heat.

Exp. Put a few drops of CS_2 at the bottom of a small glass vessel and allow them to evaporate. If a glass heated to about 300° F. be introduced, the vapor inflames. The products of combustion are CO_2 and SO_2. The analogy between CO_2 and CS_2 has been pointed out elsewhere.

Carbon Chlorides. Carbon forms with chloride at least four compounds:

C_2Cl_2	Carbon monochloride.
C_2Cl_4	" dichloride.
C_2Cl_6	" trichloride.
CCl_4	" tetrachloride.

These cannot be prepared by the direct union of their elements, but are mostly the result of the successive substitution of chlorine for hydrogen. Marsh gas, CH_4, for instance, yields, by such action, the following compounds:

CH_3Cl	Monochlorinated marsh gas.
CH_2Cl_2	Dichlorinated " "
$CHCl_3$	Trichlorinated " "
CCl_4	Tetrachlorinated " "

The third body, $CHCl_3$, is chloroform; the fourth is carbon tetrachloride, the most important of the carbon chlorides. It is a colorless, volatile liquid, which acts as a powerful anæsthetic.

Cyanogen, CN. Cyanogen is electro-negative, and in its chemical relations resembles such elements as Cl, Br and I. It forms compounds called *cyanides.* In all of these it acts as a monad; thus we have hydrogen cyanide, HCN, potassium cyanide, KCN. Dyads require, of course, two molecules of cyanogen. Calcium cyanide is CaC_2N_2 or $Ca(CN)_2$. The symbol Cy is often used in formulæ instead of the symbol CN. We write HCy instead of HCN, KCy instead of KCN, $CaCy_2$ instead of the formula for calcium cyanide just

given. The chemistry of cyanogen is more conveniently considered in connection with organic compounds.

SILICON, Si, 28.

Sources. Silicon never occurs in the free state, but is found very abundantly as oxide, SiO_2, and as silicates. It was discovered by Davy in 1807.

Preparation. By methods similar to those used for boron.

Properties. Silicon exists in three forms, amorphous, graphoidal and diamond, corresponding to those of carbon. When strongly heated in the air it burns, producing SiO_2.

Gen. Chem. Rel. Silicon is a tetrad, and is related to carbon in many ways, especially in the capacity for assuming allotropic forms. Compounds have been obtained in which it has replaced carbon. Silicon also has chemical relations to tin and titanium.

Silica, Silicic Anhydride, SiO_2. This is a widely distributed body, occurring free as common sand, chalcedony, quartz, etc., and in combination forming silicates in great variety, of which clay, granite, feldspar and sandstones are instances. A very large proportion of the solid substances in the earth's crust are forms or compounds of silica. Silica exists in the stems of grasses and in the teeth and bones of animals.

In its pure forms silicic anhydride is a colorless, nearly infusible and insoluble solid, destitute of chemical activity. In nature it often occurs in large six-sided crystals of the form shown in the cut. These are sometimes ruby-colored, and are then called amethyst. Uncrystallized silica also occurs in variously colored condi-

tions—agate, jasper, chalcedony, onyx, etc. In all its forms it is converted into a silicate by fusion with sodium carbonate, and when lime, lead oxide or other metallic oxides are mixed with the sodium silicates, we get the various forms of glass.

Silicic Acid, Orthosilicic Acid, H_4SiO_4. This cannot be produced by direct union of water with the natural forms of silica, but is obtained when silicates are decomposed by acids. If the sodium silicate be treated with hydrochloric acid, the following reaction occurs:

$$Na_4SiO_4 + 4HCl = 4NaCl + H_4SiO_4.$$

The silicic acid and salt remain in solution in the water in which the HCl was dissolved. By placing the solution in a vessel made of parchment-paper and floating it in water, the salt passes out of the vessel and leaves a pure solution of silicic acid. This process is called dialysis. Solution of silicic acid is tasteless and feebly acid to litmus. By evaporation it forms a gelatinous mass which can be brought to the composition H_2SiO_3 (metasilicic acid), and by further heating gives the insoluble anhydride. The chemistry of silicic acid is complicated. It presents some of the characters of phosphoric acid in its power of forming different acids by taking up different proportions of water. Many of the natural silicates are decomposed and rendered soluble by the combined action of frost and carbonic acid. In this way soils are formed and extensive geological changes ultimately produced.

Silicon combines with the halogens, forming bodies resembling the corresponding carbon compounds.

Silicon and fluorine have a strong affinity, and silicon fluoride, SiF_4, is easily prepared by the action of hydrofluoric acid upon silicic acid or any silicate. It is a colorless gas, which is decomposed by water, yielding gelatinous silicic acid and a double fluoride of silicon and hydrogen.

$$3SiF_4 + 4H_2O = H_4SiO_4 + 2(2HF,SiF_4).$$

The latter body is often called hydrofluosilicic acid. Silicon combines with positive elements, forming silicides, but many of these are of uncertain composition. Hydrogen silicide, H_4Si, the most interesting, is prepared by dissolving magnesium silicide in hydrochloric acid. It resembles hydrogen phosphide in taking fire spontaneously when impure. In its composition it is analogous to marsh gas, CH_4.

TIN, Sn, 118.

Sources. Tin occurs principally as dioxide, called tinstone. It was known to the ancients.

Preparation and Properties. The ore is roasted and reduced with charcoal. Tin is white, soft and easily beaten into foil, but is not tough; specific gravity, 7.28. It fuses at 442° F. (228° C.), and resists very well the action of the air and of cold acids. Nitric acid forms an insoluble dioxide. Tin forms several valuable alloys—pewter, gun-metal, type-metal, bronze and solder, elsewhere described. Speculum-metal, used for metal mirrors, is an alloy of copper and tin; glass mirrors are coated with an amalgam of tin. Tin plate is iron coated with tin by dipping it into a bath of the melted metal.

Two series of tin salts are known—*stannous*, dyad, and *stannic*, tetrad.

Stannous Oxide, SnO, is a feeble base, but is not important.

Stannous Chloride, $SnCl_2$, is formed by dissolving tin in hydrochloric acid. The solution deposits white crystals containing $2H_2O$. They dissolve in water, but are generally quickly decomposed into an oxychloride, which precipitates. Stannous chloride is an unsaturated molecule, and tends to take up chlorine or oxygen, for which reason it is used as a

reducing agent. When mixed with mercuric chloride, mercury is set free and stannic chloride formed.

$$HgCl_2 + SnCl_2 = Hg + SnCl_4.$$
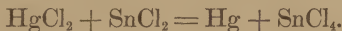
Stannous chloride is used by the dyer as a mordant under the name of tin crystals.

Stannous Sulphide, SnS, is found as a mineral, and is obtained artificially by the action of hydrogen sulphide upon stannous chloride.

$$SnCl_2 + H_2S = SnS + 2HCl.$$

Stannic Oxide, SnO_2, *Stannic Anhydride*, is found as a mineral; also produced by burning tin in the air, by oxidizing it with nitric acid and by adding an alkali to stannic chloride. Like silicic anhydride, it forms different acids according to the method of production. When tin is heated with nitric acid, a white powder is obtained, which forms complicated salts called *metastannates*. When an alkali is added to the chloride an acid is formed, having the composition H_2SnO_3, forming salts called *stannates*. This latter acid is analogous to metasilicic, H_2SiO_3, and should have been called metastannic.

Stannic Chloride, Tin Tetrachloride, $SnCl_4$, *Libavius' Fuming Liquor*. This body is largely used by dyers under the name of *nitro-muriate of tin*, being made by the action of a mixture of nitric and muriatic acid on tin. It is a colorless fuming liquid, boiling at 239.5° F. (115.3° C.).

Stannic Sulphide, SnS_2, *Mosaic Gold*, is made by passing hydrogen sulphide into stannic chloride. It is a bronze-colored powder used in printing and coloring.

TITANIUM, TI, 50, discovered by Gregor in 1709, exists as titanic anhydride, TiO_2, and also in some iron ores.

Tantalum, Ta, 182, discovered by Hatchett in 1801, and *Niobium*, Ni, 94, are found in some rare minerals.

TUNGSTEN, W, 184, discovered by Bergman in 1783, exists

as a manganese-iron tungstate, called wolfram, and as calcium tungstate (scheelite.) The only important compound of the metal is tungstic acid, H_2WO_3. Sodium tungstate has been employed for rendering dress goods fireproof.

Zirconium, Zr, 89.5, discovered by Klaproth in 1789, may belong in the aluminum group. It exists as oxide in the rare mineral *zircon.*

PLATINUM, Pt. 197.1.

Sources. Platinum occurs native, also alloyed with the elements of its class, and with gold and silver. It was discovered by Wood in 1741.

Preparation and Properties. The ore is purified by solution in *aqua regia* and precipitation of the platinum as ammonio-platinum chloride, $2AmCl + PtCl_4$. This is decomposed by heat, the platinum being left in a spongy state. This is compressed into a small bulk and hammered while red hot. It can also be fused in the flame produced by burning a mixture of oxygen and hydrogen.

Platinum is hard, white and very heavy; specific gravity, 21.5; it fuses only at a very high temperature. It resists perfectly the action of the air and of most chemical agents, and for this reason is largely used in chemical operations. It dissolves in hot *aqua regia*, forming platinum tetrachloride, $PtCl_4$. Many elements, especially zinc, tin, lead and silver, when melted with platinum dissolve it easily, forming fusible alloys. Platinum forms two series of compounds; its oxides are only feebly basic.

Platinum Tetrachloride, or *Platinic Chloride*, $PtCl_4$, is a red or brown deliquescent mass obtained by dissolving platinum in *aqua regia*, and evaporating the solution to dryness. It forms yellow granular precipitates with potassium and ammo-

nium salts, but not with those of sodium, and is of great use in analysis for the separation of potash and soda.

PALLADIUM, Pd, 106.5, discovered by Wollaston in 1803, exists associated with platinum and gold. It is hard and white, specific gravity, 11.6, and is not easily oxidized. Hammered palladium absorbs 640 times its volume of hydrogen, forming an alloy. Two series of palladium compounds are known, in which it is respectively dyad and tetrad.

RUTHENIUM, Ru, 104.4, discovered by Claus in 1846, is found with platinum, also as a sulphide. It is hard and brittle, difficult to fuse and to oxidize; specific gravity, 11.4. It acts as a dyad and tetrad.

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Calcium Group includes calcium, barium, strontium and lead. They are electro-positive dyads, not found in the free state in nature, and form oxides slightly soluble in water and much less caustic than the alkalies, but often called alkaline earths. Their sulphates, carbonates and phosphates are insoluble. They each form a dioxide which is apparently neither basic nor acid.

CALCIUM, Ca, 40.

Sources. Calcium occurs abundantly as sulphate, carbonate, phosphate and fluoride. It was discovered by Davy in 1808.

Preparation and Properties. Calcium is prepared by decomposing the chloride either by electricity or by sodium. It is light yellow, hard and malleable; oxidizes easily.

Calcium Oxide, *Quicklime*, CaO, is obtained by heating the carbonate to redness.

$$CaCO_3 = CaO + CO_2.$$

Quicklime ("quick" means alive or active) is a white,
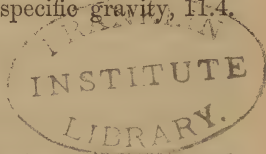
infusible solid, which neutralizes the most powerful acids and combines with water with great energy, forming

Calcium Hydrate, *Slaked Lime*, CaH_2O_2, a soft white caustic powder, slightly soluble in cold water (about 9 grains to the pint), less so in hot. The solution, known as *lime-water*, is used as a test for free carbonic acid, which produces a precipitate in it.

Slaked lime mixed with sand constitutes mortar. The cause of the hardening of mortar is not definitely known; calcium carbonate and silicate are formed, but only in small amounts.

Lime is used in agriculture to assist in decomposing the silicates in the soils.

Calcium Carbonate, $CaCO_3$, occurs abundantly and in a variety of forms. It is the chief constituent of shells and of coral. In a non-crystalline condition it is seen as *chalk*, *marble* and *limestone;* in crystals it forms *Iceland spar* and *arragonite*. It can be prepared artificially by adding sodium carbonate to calcium chloride.

$$CaCl_2 + Na_2CO_3 = CaCO_3 + 2NaCl.$$

Calcium carbonate is a white solid, insoluble in water and decomposed by a red heat. It dissolves in water containing carbon dioxide, for which reason most spring- and river-waters contain it. When present in an amount more than a grain or two to the gallon, a *hard* water is formed, which has the property of curdling soap and preventing the formation of a lather. This is due to precipitation of insoluble calcium salts, formed from the fat-acids of the soap. Boiling the water will expel the carbon dioxide and precipitate the calcium carbonate, thus softening the water. The same result is attained by adding clear lime-water, which combines with the carbon dioxide. Water which can be softened by these methods is said to be *temporarily* hard. It is probable that the calcium carbonate exists in the water as an anhydro-carbonate, anal-

ogous to borax. The excess of carbon dioxide may also be expelled by exposure of the water to air, and the calcium carbonate will then be deposited. Such an action occurs, for instance, in caves, forming *stalactites* and *stalagmites*.

Calcium Sulphate, $CaSO_4$, usually occurs as a mineral, crystallized with $2H_2O$, constituting selenite, gypsum and alabaster, sometimes, however, anhydrous. It is soluble in about 400 times its weight of cold water. It is a frequent ingredient of spring- and river-water, causing the same effect of hardness mentioned above; but as the sulphate does not owe its solubility to carbon dioxide, boiling, except for a long time, does not soften the water, and hence the condition is called *permanent* hardness. When the crystallized mineral is heated moderately, it loses its water of crystallization and becomes a soft white powder (plaster of Paris), which when mixed again with water reabsorbs it and becomes a hard mass, expanding slightly in bulk, and thus suited for taking casts of any object.

Calcium Phosphate, $Ca_3(PO_4)_2$, occurs in bone and in modified form in some mineral deposits. Its chief use is in fertilizers and in the manufacture of phosphorus and its compounds. It is insoluble in water.

Calcium Hypophosphite, $Ca(PH_2O_2)_2$, is prepared by boiling phosphorus with lime, as indicated on page 122. It is used in medicine. By substituting barium hydrate for the lime, barium hypophosphite is formed, from which *hypophosphorous acid* can be obtained as a syrupy liquid, easily decomposed and having powerful reducing action.

Calcium Chloride, $CaCl_2$, is easily obtained by dissolving the carbonate in hydrochloric acid.

$$CaCO_3 + 2HCl = CaCl_2 + H_2O + CO_2.$$

It is very soluble in water, and the solution on evaporation forms crystals of the composition $CaCl_2 + 6H_2O$, which melt and become anhydrous at about 450° F. (232° C.). The

anhydrous salt has a powerful affinity for water, and is used for drying gases. It also absorbs ammonia gas and combines with alcohol.

Calcium Hypochlorite, *Bleaching-Powder.* This body is produced by passing chlorine into slaked lime, keeping the mixture cool. The reaction should be, in theory,

$$2CaH_2O_2 + Cl_4 = Ca(ClO)_2 + CaCl_2 + 2H_2O.$$

The exact composition of the commercial bleaching-powder is, however, still undetermined; it appears to contain some unchanged calcium hydrate, but is probably mainly a combination of calcium hypochlorite, $Ca(ClO)_2$, with calcium chloride, $CaCl_2$. Bleaching-powder, when in good condition, is a loose, dry, white powder, with a faint and not disagreeable odor. If it smells of chlorine it is in bad condition. It dissolves in water. The solution possesses strong bleaching and deoxidizing powers, for which purposes it is largely used. Acids, even carbonic acid, decompose it, setting chlorine free.

$$Ca(ClO)_2, CaCl_2 + 2H_2CO_3 = 2CaCO_3 + 2H_2O + Cl_4.$$
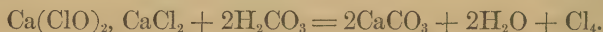
The commercial salt is often erroneously called *chloride of lime.*

Calcium Fluoride, CaF_2, *Fluor Spar*, is the principal source of the fluorine compounds. It is found as a mineral, often in very fine colored crystals, which become luminous when heated.

Tests. Calcium compounds give to flame a reddish color, which is a mixture of orange, green and faint blue. Sulphuric acid does not precipitate dilute solutions of calcium salts, as calcium sulphate is somewhat soluble in water. The usual test for calcium is ammonium oxalate, which throws down a white precipitate of calcium oxalate.

BARIUM, Ba, 137.

Sources, etc. Barium occurs principally as sulphate and carbonate. It was discovered by Davy in 1808. It is prepared from the chloride by electrical decomposition, and is a pale yellow, easily-oxidized solid, of a specific gravity of 4.

Barium Oxide, Baryta, BaO, is obtained by heating the nitrate. It easily takes water, forming barium hydrate, BaH_2O_2, which is soluble in about 20 parts of water, the solution being rendered turbid by very small amounts of carbonic acid, and hence used as a test.

Barium Dioxide, BaO_2, is formed by heating BaO in a current of air. It is used in making hydrogen dioxide.

Barium Carbonate, $BaCO_3$, is found in nature as *witherite*, and is also made artificially. It is insoluble in pure water, and used in some analytical operations.

Barium Sulphate, $BaSO_4$, *Barytes, Heavy Spar*, is found abundantly as a mineral, often forming the *gangue* or rock surrounding metallic veins. It is a very heavy, white, insoluble solid, often finely crystallized. It is used as a substitute and adulterant for white lead, and also to adulterate various other articles. When barium salts are mixed with a sulphate the barium sulphate is thrown down. Heated with carbon, it becomes a soluble sulphide, and can in this way be used as a source of the other barium salts.

Barium Nitrate, $Ba(NO_3)_2$, is used as a test solution for sulphates and in making green fire.

Barium Chloride, $BaCl_2$, is also used as a test.

Tests. Barium communicates to flame a yellowish-green color, which the spectroscope shows to contain several distinct shades.

Sulphuric acid produces in barium solutions a white precipitate of barium sulphate, insoluble in water and acids.

Strontium, Sr, 87.5, resembles barium closely in its com-

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pounds and chemical relations. It occurs as sulphate, *celestine*, and carbonate, *strontianite*. It may be prepared similarly to barium. It was discovered by Davy in 1808.

Strontium nitrate is used in making red fire.

Tests. Strontium compounds give to flame a crimson tint, which is a mixture of red, orange and blue. Its solutions produce with sulphuric acid a white precipitate resembling that given by barium. The color imparted to flame is sufficient distinction.

LEAD, Pb, 207.

Sources. Lead occurs abundantly as sulphide (*galena*), and in small quantity as carbonate, sulphate and phosphate. It was known to the ancients.

Preparation. The sulphide is the common ore. It is roasted in a free supply of air, by which a portion is converted into sulphate. The mixture is then highly heated, and the following reaction occurs:

$$PbSO_4 + PbS = Pb_2 + 2SO_2.$$

Properties. Lead is soft, bluish-white, and not capable of being made into very thin sheets or wire. It resists the action of air and of some strong acids, for which reason it is used in chemical apparatus, as in the manufacture of sulphuric acid. Pure water, free from air, has no action on lead, but aërated water soon oxidizes and dissolves it in small quantity. The presence of sulphates and phosphates interferes with this action, because they precipitate insoluble lead compounds. The composition of a water is thus a matter of some importance when the supply comes through lead pipes, and the presence of a few grains of calcium or sodium sulphate to the gallon is a good preventive of lead-poisoning. Lead melts at 617° F. (325° C.), and boils at a white heat. Specific gravity, 11.5.

It forms some important alloys. *Type-metal* contains 4 parts lead and 1 part antimony; solder, about equal parts of lead and tin; pewter, 1 part lead and 4 parts tin.

Lead Monoxide, *Litharge, Massicot,* PbO, is usually made by heating lead in the air. It is a yellowish or reddish powder, slightly soluble in water and neutralizing the most powerful acids. It fuses at a red heat, and in this condition combines easily with silica, for which reason it is often used in glazing earthenware, but such glaze is easily attacked by acids and may give rise to lead-poisoning. Lead oxide is used in paints and cements.

Lead Dioxide, Puce or Brown Oxide, PbO_2, may be obtained by the action of nitric acid upon red lead. It is a brown powder insoluble in water, and having some of the characters of an anhydride. It is an oxidizing agent.

Red Lead, *Minium,* usually Pb_3O_4, is a mixture of PbO and PbO_2. It is obtained by heating litharge in the air for some hours. It forms a bright red powder, not constant in composition. It is used as a coloring material and in the manufacture of glass.

Lead Sulphide, PbS, is abundant as a mineral, forming large cubical lead-colored crystals. Lead sulphide is easily made artificially by passing hydrogen sulphide into lead solutions; thus:

$$Pb(NO_3)_2 + H_2S = PbS + 2HNO_3.$$

Lead Carbonate, $PbCO_3$, *White Lead,* occurs as a mineral, but is now made artificially on a very large scale for use in paints. The most used method is known as the Dutch process. Thin sheets of lead are loosely rolled up and set in earthen jars, at the bottom of which some vinegar is placed. The jars are then piled in rows and covered with spent tan, but the air is not entirely excluded. The tan decomposes, producing heat and evolving carbon dioxide. A lead oxy-acetate is first formed and then converted into carbonate.

The process requires some weeks, and the product is then ground, and for use is mixed with fixed oils. Lead carbonate is a white, opaque powder, insoluble in pure water. The white lead of commerce is an oxy-salt of varying composition, approximately $2PbCO_3 + PbO + H_2O$.

Lead Sulphate, $PbSO_4$, is thrown down when lead salts are mixed with any sulphate. It is a white insoluble powder.

Lead Chloride, $PbCl_2$, is not very soluble in cold water, so that when a chloride is added to a lead salt a precipitate often occurs. It forms slender crystals, which are tolerably soluble in boiling water.

Lead Iodide, PbI_2, is easily produced by adding potassium iodide to lead solution. It forms a bright yellow powder sparingly soluble in cold water. If dissolved in boiling water, the solution on cooling will deposit the iodide in crystals.

Tests. Lead compounds give a pale green color to gas flame; with the electric spark a mixture of violet, green and yellow is produced. Sulphuric acid precipitates white lead sulphate; potassium chromate gives a yellow chromate (chrome yellow); potassium iodide, yellow lead iodide, soluble in boiling water; and hydrogen sulphide, a black lead sulphide.

Copper Group. This includes copper and mercury. They are electro-positive dyads, but also form a series of compounds in which they are apparently monad. In this latter condition they form chlorides insoluble in water, thus being connected with silver and thallium.

COPPER, Cu, 63.

Sources. Copper occurs native—*i. e.* in the free state— in large masses, also abundantly as sulphide, copper pyrites, and as oxide, silicate and carbonate. It was known to the

ancients. In small amounts it is widely distributed in nature, occurring in many articles of food, and generally in the human body, especially in the brain.

Preparation. The native copper of course requires very little preparation. The sulphide and carbonate are decomposed by roasting and successive meltings.

Properties. Copper is distinguished by its red color. It is heavy, specific gravity, 8.9; hard, and can be worked into thin plates or wire; melts at 1996° F. (1091° C.). It conducts heat and electricity very well, and resists the action of the air, but is slightly oxidized and dissolved by acids when in contact with air. Even sea-water and the acids of fruits will produce this effect, and hence the danger of using copper vessels for kitchen purposes. It furnishes some valuable alloys —brass, gun-metal, etc. An alloy of copper with hydrogen, Cu_2H_2, has attracted some attention on account of its supposed theoretical relations. Copper compounds form two sets of salts; in the most important, the *cupric* series, the metal is dyad; in the other, *cuprous*, two atoms of copper are supposed to act together as a double atom. The condition is usually explained by graphic formulæ (see page 31), thus:

Copper in cupric salts,   Copper in cuprous salts,
—Cu—   —Cu—Cu—

The cuprous salts are mostly colorless; the cupric salts are green or blue.

Copper Monoxide, Black Oxide, CuO, is prepared by heating copper in air or by roasting the nitrate. It is a heavy black powder, dissolving in acids. Heated in a current of hydrogen it is easily reduced. Copper oxide is used in organic analysis. Copper hydrate, CuH_2O_2, formed when copper salts are mixed with an alkali, is a bluish-green mass, dissolving in ammonia, producing a clear, deep-blue liquid of complicated composition. With potassa and soda no solution occurs except in the presence of certain organic bodies, especially sugar, when a

14 *

clear blue solution is also formed. If such solution is boiled, the cupric hydrate is changed to cuprous, which is precipitated as a red or orange powder. This reaction is a useful test for sugar (*q. v.*).

Copper Sulphide, CuS, *Copper Pyrites*, occurs as a mineral, and is easily produced as a black precipitate by mixing copper solution with hydrogen sulphide.

$$CuSO_4 + H_2S = CuS + H_2SO_4.$$

Copper Carbonate, $CuCO_3$, appears not to be known in the pure state. Various oxycarbonates, malachite and azurite, exist as fine minerals, and similar compounds are obtained by the addition of carbonates to copper salts. The natural forms are used for ornamental articles, the artificial form for paints.

Copper Sulphate, $CuSO_4$, *Blue Vitriol*, *Blue Stone*, is formed by dissolving copper or its oxide in sulphuric acid. It forms large blue crystals soluble in water, and having the composition $CuSO_4 + 5H_2O$. When heated, the water of crystallization is driven out and the salt becomes a soft white powder.

Copper Nitrate, $Cu(NO_3)_2$, is a residue from the preparation of nitric oxide. It has little importance.

Copper Chloride, $CuCl_2$, is in green crystals, soluble in water. It forms a number of double salts. An oxychloride is used as a paint.

Copper Arsenite, Scheele's or Paris Green, $CuHAsO_3$, is a bright green powder, obtained by mixing an alkaline arsenite with cupric sulphate. It is used for killing potato-bugs and also as a color. It is a violent poison. A compound of acetate and arsenite is known as Schweinfurth green.

Cuprous Salts. These are of little importance. *Cuprous Oxide*, Cu_2O, has already been mentioned as the result of the action of sugar on a mixture of caustic alkali and cupric hydrate; *Cuprous Chloride*, Cu_2Cl_2, is a white solid, insol-

uble in water. The cuprous salts are easily converted into cupric.

Tests. Copper gives a green tint to flame; with the electric spark it gives a mixture of violet and green light. Ammonia gives a bluish-green precipitate; and on adding it in excess a deep blue color. Potassium ferrocyanide gives a mahogany brown precipitate of copper ferrocyanide. Hydrogen sulphide gives a brown precipitate of copper sulphide. A clean piece of iron immersed in a solution of copper becomes quickly covered with a bright red coating of copper.

MERCURY, Hg, 200.

Sources. Mercury is found native, and as the sulphide (cinnabar) in Spain, California, Japan and China. It was known to the ancients.

Preparation. The cinnabar is roasted; the sulphur is evolved as sulphur dioxide (SO_2), the mercury volatilizes, and the vapor is condensed in earthen pipes.

Properties. Liquid at the ordinary temperature, freezing at —40° F. and C., and boiling at 675° F. (357° C.); when pure it does not tarnish in dry or moist air, but above 300° C. it absorbs oxygen and passes into the red oxide. It is largely employed in the processes for extracting silver and gold from their ores, and is used in medicine. It is very lustrous and heavy; specific gravity, 13.56. Mercury has the power to dissolve other elements, forming alloys which are known as amalgams. These are either soft or hard according to the quantity of mercury used. Mirrors are coated with an amalgam of tin. Dentists use extensively various amalgams for filling teeth. When mercury is triturated with a soft substance, it can be so finely divided as to lose all its lustre and appear as a bluish or gray powder, which is used in medicine under the name " blue pill." The vapor is 100 times as heavy

as hydrogen—half as heavy as theory would require. Two series of salts are known, corresponding to the copper salts, and called respectively *mercurous* and *mercuric* salts.

Mercuric Oxide, HgO, *Red Precipitate*, can be obtained by keeping mercury for some time at its boiling-point, also by heating mercuric nitrate. It is a red or yellowish-red powder, decomposed at a dull red heat. It is an active base, and dissolves very slightly in water. It is used in medicine. When caustic soda is added to a mercuric solution, mercuric hydrate, HgH_2O_2, is thrown down as a yellow precipitate.

Mercuric Sulphate, $HgSO_4$, is formed by boiling mercury with sulphuric acid. The reaction is

$$Hg + 2H_2SO_4 = HgSO_4 + 2H_2O + SO_2,$$

similar to that with copper. Mercuric sulphate is a white powder, which is decomposed by water, forming a yellow oxysulphate, $HgSO_4 + 2HgO$, called turpeth mineral.

Mercuric Nitrate, $Hg(NO_3)_2$, is generally seen in solution with excess of nitric acid, forming the *acid mercury nitrate* used in medicine. The normal salt is used as a source of the oxide.

Mercuric Chloride, $HgCl_2$, *Corrosive Sublimate*, is obtained by heating a mixture of mercuric sulphate and common salt.

$$HgSO_4 + 2NaCl = HgCl_2 + Na_2SO_4.$$

The mixture is strongly heated; the mercuric chloride rises in vapor and condenses on a cool surface. This process is called *sublimation*. Corrosive sublimate is a heavy, white, crystalline powder, soluble in water and ether, and having an acrid metallic taste. It is extremely poisonous, about five grains being a fatal dose. It forms with albumen an insoluble precipitate.

Mercuric Iodide, HgI_2, *Red Iodide*, is formed by mixing corrosive sublimate with potassium iodide.

$$HgCl_2 + 2KI = HgI_2 + 2KCl.$$
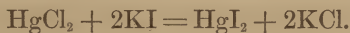
Mercuric iodide is at first yellow, but rapidly changes to a brilliant scarlet. It is soluble both in $HgCl_2$ and KI, and the solutions are used as tests for ammonia and various organic principles.

Mercuric Sulphide, HgS, *Vermilion*, is found as a red mineral, *cinnabar*, an important ore of mercury. It can be prepared by passing hydrogen sulphide into mercury chloride.

$$HgCl_2 + H_2S = HgS + 2HCl.$$

Thus prepared it is black, but is converted into the red form by heating. It is used as a bright red paint.

The *Mercurous Salts* are mostly of little importance.

Mercurous Oxide, Hg_2O, is a black powder, easily decomposed, obtained by the action of alkaline hydrates upon calomel.

Mercurous Chloride, Hg_2Cl_2, *Calomel*, is formed similarly to corrosive sublimate, substituting mercurous sulphate for mercuric.

$$Hg_2SO_4 + 2NaCl = Hg_2Cl_2 + Na_2SO_4.$$

The calomel is driven out by heat, and forms a sublimate. It is a white, heavy, tasteless powder, insoluble in water.

Mercurous Sulphate, Hg_2SO_4, may be made by rubbing mercuric sulphate with another atomic proportion of mercury.

$$Hg + HgSO_4 = Hg_2SO_4.$$
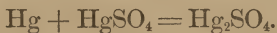
Tests. Mercury imparts no characteristic color to flame. The precipitation tests are different for the two series of salts. Mercurous salts give with hydrochloric acid a white precipitate which is blackened by ammonia. Mercuric salts give with potassium iodide a yellow precipitate of HgI_2, changing to scarlet and soluble in excess of the precipitant.

Any compound containing mercury will give with Reinsch's test (*q. v.*) a bright silvery coating on copper foil; also, with hydrogen sulphide in excess a black precipitate.

Zinc Group. This includes zinc, magnesium, cadmium and beryllium. They are never found free, but are tolerably easily reduced from their compounds. They each form but one definite oxide, which is insoluble in water, not caustic, but capable of forming well-marked salts. Beryllium is somewhat uncertain in its relations.

ZINC, Zn, 65.

Sources. Zinc exists rather abundantly as sulphide (blende), carbonate (calamine), silicate (electric calamine) and as oxide. Zinc was known in the thirteenth century.

Preparation. The ores are first converted into oxide by roasting in the air, and the oxide is then heated with charcoal in an earthen retort or crucible so arranged that the zinc may distil into a receiver.

Properties. Zinc is hard, bluish-white, generally decidedly crystalline. Sp. gr. 7.14. It melts at 770° F. (410° C.), and distils at about a red heat. It is brittle at ordinary temperatures, but becomes soft and malleable at between 212° and 302° F. (100° and 150° C.), and at a higher temperature again becomes brittle. Zinc oxidizes slightly in moist air, and when highly heated burns with a greenish-white flame, producing ZnO. Acids and strong alkalies dissolve it. The uses of it are numerous. It is employed in making several important alloys, as brass and gun-metal, which contain copper and zinc, and German silver, which contains copper, zinc and nickel. Galvanized iron is simply iron covered with a layer of zinc by dipping it in a bath of melted zinc. Commercial zinc is very likely to contain arsenic.

Zinc Oxide, ZnO, *Zinc White*, is easily made by burning zinc. It is a soft powder, yellow when hot, white when cold. It is used as a paint, as an application in surgical dressings and as a face-powder. It dissolves easily in acids.

Zinc Hydrate, ZnH_2O_2, is produced by adding ammonia to a solution of a zinc salt. It is a white precipitate, soluble in acids and alkalies.

Zinc Sulphide, ZnS, occurs as a crystalline mineral called *blende*. It is produced artificially as a white gelatinous mass by action of hydrogen sulphide on alkaline solutions of zinc.

Zinc Carbonate, $ZnCO_3$, *Calamine*, exists as a mineral. When zinc sulphate and sodium carbonate are mixed, an oxy-salt,

$$3ZnCO_3 + 5ZnO + 6H_2O,$$
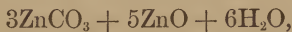
is thrown down.

Zinc Sulphate, $ZnSO_4$, *White Vitriol*, is made either by dissolving zinc in sulphuric acid or by oxidizing the sulphide. It forms white crystals having the formula $ZnSO_4 + 7H_2O$. They are soluble in water, have an acid reaction and act as an emetic.

Zinc Chloride, $ZnCl_2$, is made by dissolving zinc scraps in hydrochloric acid. It forms white masses, which absorb water rapidly from the air (*deliquesce*) and make a strong solution. Zinc chloride is a powerful corrosive, coagulates albuminous matter, and is used as a preservative in anatomical preparations, also as an application in dentistry. When a strong solution of zinc chloride is mixed with zinc oxide, the two combine and form a hard, white insoluble mass which is used as a filling for teeth.

Zinc Phosphate, or rather *Oxy-phosphate*, made by mixing zinc oxide with phosphoric acid, and consisting of zinc phosphate united with zinc oxide, has come into use lately as a substitute for the oxychloride in filling teeth.

Tests. Zinc burns with a flame which is a mixture of red

and blue. The best liquid test depends on the fact that a white precipitate is thrown down by the action of hydrogen sulphide on alkaline solution of zinc salts.

MAGNESIUM, Mg, 24.

Sources. Magnesium occurs as carbonate (dolomite), silicate (talc and soapstone), also as hydrate and chloride. Most natural water contains magnesium compounds. It was discovered by Davy in 1808.

Preparation and Properties. It is obtained by heating the chloride with sodium. It is bright and malleable. Sp. gr. 1.74. When strongly heated in the air it burns with a very bright light, producing MgO. Magnesium compounds often cause *hardness* in water similar to that produced by calcium salts.

Magnesium Oxide, MgO, *Magnesia*, is usually obtained by heating the carbonate to redness. It is a light, white powder, very feebly soluble in water, and neutralizing acids. Magnesium hydrate is also known.

Magnesium Carbonate, $MgCO_3$, occurs as a mineral. It is usually obtained by adding sodium carbonate to magnesium sulphate. In this case an oxycarbonate is formed, which is known as *magnesia alba*.

Magnesium Sulphate, $MgSO_4$, *Epsom Salt*. This body is obtained on the large scale by the action of sulphuric acid upon natural magnesium carbonate and silicate. It forms colorless crystals, *isomorphous* (see page 172), with zinc sulphate. As thus crystallized the salt has the composition $MgSO_4 + 7H_2O$. It is very soluble in water.

Magnesium Chloride, $MgCl_2$, resembles zinc chloride in some respects, but does not have the same corrosive action.

Tests. Magnesium gives no color to a gas flame; with the electric spark it gives a green light. A mixture of ammonia, ammonium chloride and sodium phosphate is used as a test for magnesium, but it is not applicable until after all bodies except compounds of the potassium group are removed from the solution.

CADMIUM, Cd, 112, occurs in various zinc ores. It is extracted in the same manner as zinc, and is silver-white and crystalline. Sp. gr. 8.6. It melts at 442° F. (228° C.), and is nearly as volatile as mercury. It is easily dissolved by ordinary acids. Its salts are of little interest, except the sulphide. CdS is obtained as an orange-yellow precipitate by passing hydrogen sulphide into a solution of cadmium.

Beryllium, or *Glucinum*, Be, 9.4, occurs in beryl, emerald and a few other minerals. It is white and malleable. Specific gravity, 2.1.

Thorium, Th, 231.5, was discovered by Berzelius in 1828 in a silicate called *thorite.*

Yttrium, Y, 92, was discovered by Gadolin in 1794, and *Erbium*, E, 169, was discovered by Mosander in 1843. They exist in a mineral called *gadolinite,* found at Ytterby, Sweden.

The last three elements have been so little studied on account of their rarity that it is as yet not possible to classify them definitely.

Iron Group. This includes aluminum, iron, manganese, nickel, cobalt and partly chromium. They are not found native except in small quantity. They form two sets of compounds, acting in one set as dyads and in the other as double tetrads. Several are capable of forming well-marked anhydrides. Uranium, iridium and rhodium are also described in this connection, although their exact relations are not made out.

ALUMINUM, Al, 27.5.

Sources. Aluminum is very abundant as silicate, constituting clay and many common rocks. Most building materials are mixtures of aluminum silicate with other silicates. It was discovered by Wohler in 1828.

Preparation. By decomposing the chloride with sodium and by several other methods.

Properties. Aluminum is white and not very lustrous, malleable and ductile, sonorous and very light; specific gravity, 2.6. It tarnishes slightly in the air, and dissolves rapidly in hydrochloric acid and in caustic alkalies, but not in nitric acid. It melts at 842° F. (450° C.). Its alloys are very valuable; that with copper, aluminum bronze, has the color of gold and keeps well in the air.

Aluminum forms but one series of compounds, which possess strong analogies to the tetrad series of iron salts; hence aluminum is regarded also as a tetrad. Like iron, it forms compounds by the joint action of two tetrad atoms, which act as a hexad, as explained on pages 29 and 31.

Aluminum Oxide, Al_2O_3, occurs naturally as *corundum*, which, when crushed, constitutes *emery;* finely crystallized, as the *ruby* and *sapphire*. It can be prepared by heating ammonia alum; it then forms a white powder. In all the anhydrous forms it is absolutely insoluble in water, and almost so in acids and alkalies.

Aluminum Hydrate, $Al_2H_6O_6$, is usually obtained by adding ammonia to a solution of the sulphate.

$$Al_2(SO_4)_3 + 6AmHO = 3Am_2SO_4 + Al_2H_6O_6.$$

This hydrate is a gelatinous white mass, easily soluble in acids and alkalies, and has a strong affinity for organic matter; with organic colors it forms precipitates called lakes. In the art of dyeing, alumina is used as a *mor-*

dant—that is, a material to cause the color to adhere firmly to the fabric.

Aluminum oxide appears to possess a property similar to phosphoric anhydride, P_2O_5, of forming hydrates with different proportions of water:

$$Al_2O_3 + H_2O = Al_2H_2O_4 \qquad \text{Metalumina.}$$
$$Al_2O_3 + 3H_2O = Al_2H_6O_6 \qquad \text{Orthalumina.}$$
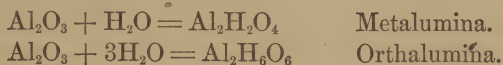
These are found as minerals, the first known as diaspore, the second as gibbsite.

Aluminum Sulphate, $Al_2(SO_4)_3$, is now largely manufactured for use in dyeing and in other operations. It is prepared either by the action of sulphuric acid upon clay nearly free from iron, or by the roasting of a mineral termed *alum shale*, which contains iron disulphide, FeS_2, and aluminum silicate. The roasting converts the sulphur into sulphuric acid, which then forms aluminum sulphate and iron sulphate. These are dissolved in water and separated by crystallization. $Al_2(SO_4)_3$ usually forms a white crystalline mass having an acid reaction.

Alum. The word "alum," like the term "alcohol," long applied to a single substance, has now come to refer to a group of substances characterized by similarity of composition and chemical relations. The alums are a series of double sulphates. One kind of alum has the formula,

$$Al_2(SO_4)_3 + K_2SO_4 + 24H_2O.$$

The aluminum in this compound may be replaced by most of the elements of this group. The potassium may be replaced by any element of its group, or by ammonium, giving a series of salts of which a few instances are appended:

$Al_2(SO_4)_3 + K_2SO_4 + 24H_2O$	Potassium alum.
$Cr_2(SO_4)_3 + Na_2SO_4 + 24H_2O$	Sodio-chromic alum.
$Fe_2(SO_4)_3 + Am_2SO_4 + 24H_2O$	Ammonio-ferric alum.
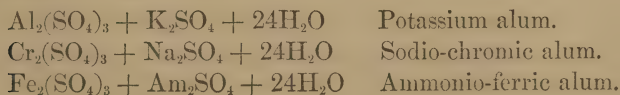
The particular point about these compounds is that they all contain the same amount of water of crystallization and all crystallize in octahedra. Bodies that are similar both in composition and in crystalline form are called ISOMORPHOUS.

Common Alum, Ammonia Alum, crystallizes in large clear octahedra, which dissolve easily in water, the solution being acid to test-paper and strongly astringent.

When alum crystals are gently heated they swell up, lose their water of crystallization and fall to a soft white powder— burnt alum.

Aluminum Chloride, Al_2Cl_6, is prepared by heating alumina and charcoal in a current of chlorine.

$$Al_2O_3 + 3C + 6Cl = 3CO + Al_2Cl_6.$$
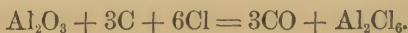
Glass, Pottery and Porcelain.—These important articles are mixtures of various silicates, especially aluminum, calcium and sodium silicates.

Colored glasses are produced by the use of various oxides. Ferrous oxide, FeO, produces a deep green (bottle glass), and manganese dioxide communicates a purple tint. It is difficult to obtain materials entirely free from iron which give a green tinge; for a colorless glass manganese dioxide is added.

Pottery and earthenware are made of clay, moulded while wet and then baked at a temperature which renders it no longer plastic. The glaze is a fusible sodium silicate made from a mixture of sand and salt. Lead silicate is also used. Porcelain is a mixture of feldspar (aluminum and potassium silicate), sand and *kaolin*, a hydrated aluminum silicate. These, when mixed with water, moulded to shape and heated to a high temperature, form a porous ware called *biscuit*. This is glazed by fusing on the surface various silicates.

The colors on pottery and porcelain are mostly mixtures of sand and metallic oxides.

Aluminum silicate exists in combination with other silicates in a large number of minerals.

Tests. Aluminum compounds give no color to common flames. When intensely heated by electric sparks a number of tints, chiefly red and blue, are produced. The precipitation tests are not always satisfactory. Ammonium hydrate throws down a gelatinous precipitate of aluminum hydrate.

IRON, Fe, 56.

Sources. Iron occurs very abundantly as oxide, sulphide and carbonate; some of its compounds occur in small quantities in many rocks and soils, and hence it is dissolved by natural waters and taken up by plants; from these sources it gets into the animal body, and is an essential constituent of the blood of the higher animals. Iron was known to the ancients. It sometimes occurs native, especially in meteoric stones.

Preparation and Properties. The preparation of iron is a simple process of reduction. A fine-grade Quevenne's iron, for use in medicine, is made by passing hydrogen over sesquioxide. On the large scale, iron ore, which generally consists of an oxide or carbonate mixed with clay, sand and other minerals, is heated in blast-furnaces with coal and limestone. The limestone makes a fusible calcium silicate, *slag;* the coal takes the oxygen away from the iron. The melted mass is then run out into thick bars, forming *pig* or *cast iron*, the most impure form, containing four or more per cent. of carbon, also phosphorus, sulphur, silicon and other bodies. The arrangement of a blast-furnace is shown in section in the cut. The mixture of ore, limestone and coal is put in at the top, and as it descends the chemical changes occur and the melted iron collects in the lower part or *crucible, c, d, e;* the

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slag floats on the iron, and is drawn off from time to time. The iron is run out by the opening between g and f. A blast of hot air is introduced at d.
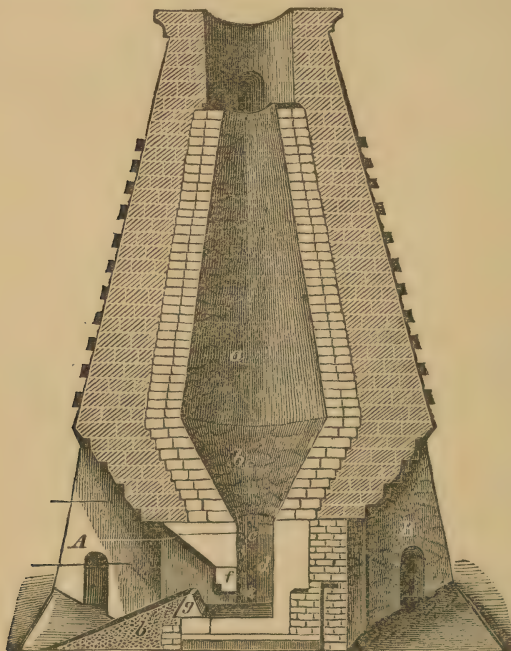
This crude iron is worked in a furnace in a current of air, by which the impurities, especially the carbon, are burnt out,

and the iron becomes less fusible and much tougher, constituting *wrought iron*. A section of the furnace is shown in the cut. The fuel is placed at a, and the material to be heated at d, d. The flame is thrown down on the iron by the arched roof, for which reason it is called a reverberatory furnace.
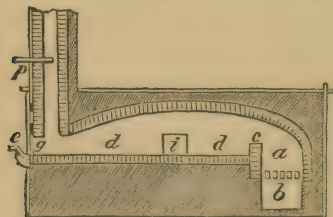
Steel contains about one-half per. cent of carbon; it is therefore intermediate in composition. Steel was formerly made by heating wrought iron with charcoal; it is now made directly from cast iron by burning out a certain percentage of the carbon (Bessemer process). Pure iron is very soft, but is found in commerce with various impurities which give hardness and other special qualities to the iron; carbon gives hardness and fusibility; phosphorus and sulphur give fusibility and great brittleness, and are very objectionable. Cast iron melts at about 3000° F. (1649° C.). Iron is strongly magnetic and not much affected by dry air, but is oxidized by moist air and easily dissolved by acids. It forms two series of salts—*ferrous*, in which it is a dyad, and *ferric*, in which it is apparently a triad, but the formulæ of the ferric compounds are generally doubled, and the iron is regarded as a double tetrad acting as a hexad.

$$\text{Dyad iron,} \qquad\qquad \text{Hexad iron,}$$
$$\text{—Fe—} \qquad\qquad\quad \equiv\text{Fe—Fe}\equiv$$

Ferrous salts are generally green; ferric salts brown or red.

Conversion of one Series of Iron Salts into the other.

Ferrous salts are converted into ferric by oxidizing agents. As the ferric salts are written with double formulæ, the reaction will always require two molecules of ferrous for one of ferric; thus:

$$2FeO + O = Fe_2O_3.$$
$$2FeCO_3 + O = Fe_2O_3 + 2CO_2.$$
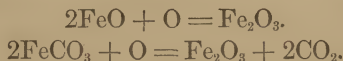
As nitric acid gives up three atoms of oxygen as an oxidizing agent (page 116), it follows that in writing reactions in which it is used six molecules of the ferrous salt must be taken.

To make a normal ferric salt we must add one-half as much of the negative element as the ferrous salt already contains;

that is, one molecule of the radicle for every two molecules of the ferrous salt. For instance, in making ferric chloride, if we take $6FeCl_2 + 2HNO_3$, we will get an oxychloride or oxynitrate. The complete reaction is

$$6FeCl_2 + 6HCl + 2HNO_3 = 3Fe_2Cl_6 + 4H_2O + 2NO.$$
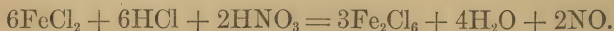
Ferrous salts are formed from ferric by the action of reducing agents, especially nascent hydrogen or sulphurous acid. The nascent hydrogen is usually obtained by adding to the ferric salt a mixture of zinc and sulphuric acid. With ferric chloride the reaction would be

$$Fe_2Cl_6 + H_2 = 2FeCl_2 + 2HCl.$$
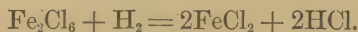
With ferric sulphate we should have

$$Fe_2(SO_4)_3 + H_2 = 2FeSO_4 + H_2SO_4.$$

Hydrogen sulphide will also reduce ferric salts:

$$Fe_2Cl_6 + H_2S = 2FeCl_2 + 2HCl + S.$$

Ferrous Oxide, FeO, is difficult to obtain. Ferrous hydrate, FeH_2O_2, is formed as a white precipitate by the action of caustic alkali upon a ferrous salt. It immediately begins to change by absorbing oxygen, and becomes ferric oxide.

Ferric Oxide, Fe_2O_3, *Red Oxide, Sesquioxide*, occurs frequently in small amounts in many minerals, and also as iron ore, called red hematite or specular iron. It may be prepared artificially by heating ferrous sulphate ($2FeSO_4 = Fe_2O_3 + SO_2 + SO_3$), and is the residue obtained in making Nordhausen sulphuric acid. It is a soft red powder, difficult to dissolve in acids. The finer grades constitute *rouge;* the coarser, *venetian red* and *crocus*, are used for paints.

Ferric Hydrate, $Fe_2H_6O_6$, is easily formed by adding caustic alkali to a ferric salt.

$$Fe_2Cl_6 + 6AmHO = 6AmCl + Fe_2H_6O_6.$$

Ferric hydrate is a soft brown mass, insoluble in water, but

dissolving easily in acid. Its chief importance is as an antidote to arsenic, for which use it must be freshly prepared. Ordinary iron rust consists of impure ferric hydrate, and it also occurs in an impure condition as brown hematite, a very important ore of iron.

Magnetic Iron Oxide, FeO,Fe_2O_3, a union of the two oxides, is found as a finely crystallized mineral and valuable ore of iron. It can retain magnetism, and is occasionally found in a magnetized condition, constituting *loadstone*.

Ferrous Sulphide, FeS, made by fusing iron with sulphur, is a dark slag-like mass, used as a source of hydrogen sulphide.

Iron Disulphide, FeS_2, *Iron Pyrites*, is abundant as a mineral, crystallized in brass-colored cubes often mistaken for gold, and hence called fool's gold. It is of no use as an iron ore, on account of the sulphur, but is used as a source of sulphuric acid and some sulphates.

Ferrous Carbonate is a mineral and valuable iron ore, and exists in many mineral (chalybeate) waters. It is produced by mixing ferrous sulphate with sodium carbonate.

$$FeSO_4 + Na_2CO_3 = FeCO_3 + Na_2SO_4.$$

In this form, however, and also as dissolved in water, it is prone to oxidation, passing into the condition of ferric hydrate, which forms a red deposit. This oxidation is hindered by sugar, and a ferrous carbonate, known as Vallet's mass, is prepared for medical use by precipitating it in contact with sugar. The natural carbonate is permanent in the air.

Ferric Carbonate has not been obtained.

Ferrous Sulphate, $FeSO_4$, *Green Vitriol*, *Copperas*, is formed by dissolving iron in sulphuric acid or by oxidizing iron pyrites. It forms clear green crystals containing $FeSO_4 + 7H_2O$, easily soluble in water, the solution being liable to oxidation.

It is used in the manufacture of ink and in dyeing. The crystals effloresce on exposure to air. Ferrous sulphate is a reducing agent, and is called a disinfectant, but merely acts as a deodorizer by absorbing some sulphur compounds. It forms double salts with ammonium and potassium sulphates.

Ferric Sulphate, $Fe_2(SO_4)_3$, is made by heating together a mixture of ferrous sulphate, sulphuric acid and nitric acid. The reaction is

$$6FeSO_4 + 3H_2SO_4 + 2HNO_3 = 3Fe_2(SO_4)_3 + 4H_2O + 2NO.$$

If one-half the above quantity of sulphuric acid be used, an oxysulphate, $5Fe_2(SO_3)_3 + Fe_2O_3$, is formed, called *Monsel's Salt*, or when dissolved *Monsel's Solution*, and used as a styptic.

Ferric Chloride, Fe_2Cl_6, is generally seen as an alcoholic solution called muriated tincture of iron. Ferric chloride is made either by dissolving ferric hydrate in hydrochloric acid,

$$Fe_2H_6O_6 + 6HCl = Fe_2Cl_6 + 6H_2O,$$

or by boiling ferrous chloride with nitric and hydrochloric acid,

$$6FeCl_2 + 6HCl + 2HNO_3 = 3Fe_2Cl_6 + 4H_2O + 2NO.$$
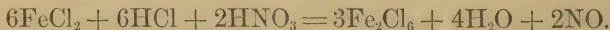
When the solution is evaporated a red crystalline mass of $Fe_2Cl_6 + 6H_2O$ is formed, which is decomposed by heat. The anhydrous salt is yellow, and extremely deliquescent. The solution dissolves ferric hydrate, forming a ferric oxychloride.

Iron nitrates, iodides and bromides do not possess special interest.

Tests. Iron gives no color to gas flames; with the electric spark it gives a mixture of many tints. Ferrous salts are characterized by their green color, ferric by red or brown. The following are the principal liquid tests:

	With Ferrous Salts.	With Ferric Salts.
Ammonia,	Green ferrous hydrate, turning red,	Red ferric hydrate.
Potassium ferrocyanide,	Light blue precipitate,	Dark blue precipitate (Prussian blue).
Potassium ferricyanide,	Dark blue precipitate (Turnbull's blue),	No precipitate.
Tannin,	No action,	Black precipitate.
Potassium sulphocyanate,	No action,	Blood-red color, but no precipitate.

MANGANESE, Mn, 55.

Sources. Manganese exists principally as oxide, also as sulphide, carbonate and silicate. It was discovered by Gahn in 1780.

Preparation and Properties. By reducing the oxide or carbonate with charcoal, or by electrical decomposition of the chloride. It is grayish-white, brittle and hard; specific gravity between 7 and 8. It forms two series of salts parallel to those of iron.

Manganous Oxide, MnO, is obtained by igniting the carbonate; the *hydrate*, MnH_2O_2, is obtained by adding an alkali to a manganous salt. Both substances absorb oxygen rapidly.

Manganous Chloride, $MnCl_2$, may be obtained from the residue left after the preparation of chlorine. It forms pink crystals, deliquescent and soluble in water.

Manganous Sulphate, $MnSO_4$, is made by dissolving the oxide or carbonate in sulphuric acid. It is a rose-colored salt, soluble in water, and used in dyeing. It crystallizes with $7H_2O$.

Manganese Dioxide, MnO_2, *Black Oxide.* This is rather an abundant mineral, occurring crystallized or in masses. It is extensively used as an oxidizing agent and in the manufacture of chlorine. Ordinarily it is in black masses or powder. It conducts electricity and dissolves in acids with evolution of oxygen. When heated to redness it is converted into an intermediate oxide, Mn_3O_4.

Manganic Oxide, Mn_2O_3, is found as a mineral. It is a weak base.

MANGANESE ACIDS. Manganese is capable of forming two anhydrides, which are not known in the free state, but of which some of the salts are common. The acids are—

$$H_2MnO_4 \qquad \text{Manganic acid.}$$
$$H_2Mn_2O_8 \qquad \text{Permanganic acid.}$$

The former has not been obtained.

Manganates. These are formed by fusing manganese dioxide with caustic alkali and potassium nitrate or chlorate. In this way potassium manganate, K_2MnO_4, is formed as a green crystalline mass. When dissolved in water the manganates turn into permanganates by the following reaction:

$$3K_2MnO_4 + 2H_2O = K_2Mn_2O_8 + MnO_2 + 4KHO.$$

The change of composition is indicated by a change of color from green to red, for which reason the potassium manganate was called *chameleon mineral.*

Potassium Permanganate, $K_2Mn_2O_8$, is now made in large amount as an oxidizing and deodorizing agent. The solution slowly decomposes when exposed to the air, and is supposed to give off ozone, and thus act as a disinfectant. It is decomposed by organic matters, by sulphites and sulphides and reducing agents generally, becoming converted into a colorless solution. It can therefore be employed not only to destroy organic matter, but also as a measure of the amount present.

Tests. Manganese compounds give to gas flame a mixture of yellow and green light. The principal test is the flesh-colored precipitate, MnS, produced by ammonium sulphide.

NICKEL, Ni, 59.

Sources. Nickel occurs principally in union with arsenic and sulphur; also in meteoric iron as an alloy. It was discovered by Cronstedt in 1751.

Preparation. Nickel is reduced by roasting and reduction with charcoal.

Properties. It is hard and white, of specific gravity 8.8, fusing at a high temperature and resisting the action of air at common temperatures. Like iron, it can acquire permanent magnetism. Solution of nickel can be decomposed by an electric current, and nickel-plating is performed in this way. An alloy of copper, zinc and nickel is called German silver, and some other alloys have been used in coins.

Nickel Monoxide, NiO, is obtained by heating the nitrate or carbonate. A hydrate, NiH_2O_2, is obtained in the usual manner. It is green and soluble in ammonia. Its compounds with acid are mostly bright green.

Nickel Sesquioxide, Ni_2O_3, is also known, but does not appear to form salts.

Nickel Sulphate, $NiSO_4$, is the most important salt. It usually crystallizes with 7 molecules of water, and can easily form double salts with potassium and ammonium sulphates.

Tests. Nickel compounds give no color to flame. They are characterized by bright green or blue color, and a green precipitate with potassium cyanide.

Cobalt, Co, 59, discovered by Brandt in 1733, is found associated with nickel, which it closely resembles in properties and chemical relations. Its compounds are mostly red or

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blue. The element itself is hard, white, magnetic and difficult to fuse; specific gravity, 8.7. The oxides, sulphates, carbonates, etc. resemble in composition those of nickel. It produces with the electric spark a mixture of colors. The tests are nearly the same as those of nickel, the distinction being that potassium cyanide forms with cobalt salts a compound which is not so easily decomposed by acids as the corresponding nickel compound.

CHROMIUM, Cr. 52.2.

Sources. Chromium occurs principally as an oxide in combination with iron oxide, constituting chrome iron ore, FeO,Cr_2O_3; also as lead chromate, $PbCrO_4$. It was discovered by Vauquelin in 1797.

Preparation and Properties. By heating the oxide with charcoal, chromium is obtained as a hard crystalline mass, not easily oxidized or dissolved. It forms two sets of salts, analogous to those of iron, and also a marked anhydride, which forms salts isomorphous with the sulphates. By this fact chromium is partly related to the oxygen group. The compounds in which it acts as a positive metal are of very little importance; they agree mostly with the corresponding iron and manganese compounds. The chromous salts are unstable. Almost all the chromium compounds are high-colored.

Chromic Oxide, Chromium Sesquioxide, Cr_2O_3, *Chrome Green,* obtained by decomposing some of the chromates (*q. v.*), is a bright green powder used as a paint.

Chromic Anhydride, CrO_3, is easily obtained by the action of acids upon the chromates. It forms bright red crystals, very deliquescent, soluble in water and having powerful oxidizing properties.

Chromates are formed by heating chrome iron ore with alkali and nitre. The most common salts are those of potassium.

Potassium Chromate, K_2CrO_4, forms lemon-yellow crystals soluble in water.

Potassium Anhydrochromate, $K_2CrO_4CrO_3$, commonly known as *bichromate,* is in large, bright red crystals which are soluble in water. It is extensively used in dyeing and as a source of various colors.

Lead Chromate, Chrome Yellow, $PbCrO_4$, is easily formed by adding a soluble chromate to a lead salt.

$$Pb(NO_3)_2 + K_2CrO_4 = PbCrO_4 + 2KNO_3.$$
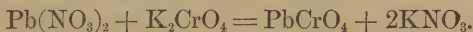
It is bright yellow and insoluble in water. An oxychromate, $PbCrO_4 + PbO$, is known as a scarlet pigment.

Chromates are decomposed when heated with organic matter, especially in the presence of an acid. The change of composition is generally indicated by a change of color from the yellow or red of the chromate to the green of the chromic salt.

Exp. Add to a dilute solution of potassium anhydrochromate a few drops of alcohol and some hydrochloric acid, and boil the mixture for a few moments. Vapors of aldehyde (*q. v.*) will be given off, and the red color of the liquid will change to green, chromic chloride being formed.

The reaction is,

$$K_2CrO_4CrO_3 + 8HCl = Cr_2Cl_6 + 4H_2O + 2KCl + O_3.$$

The O_3 oxidizes the alcohol to aldehyde.

A mixture of potassium anhydrochromate and sulphuric acid is used as an oxidizing agent in galvanic batteries. The chromic acid becomes reduced to sesquioxide, and forms chromic sulphate; the liquid turns green, and afterward deposits dark ruby-red crystals of chrome-alum.

$$K_2SO_4 + Cr_2(SO_4)_3 + 24H_2O.$$

Tests. Chromium compounds communicate no color to flame; with the electric spark they give a mixture of green and dark blue. Chromous and chromic compounds are not often encountered in analysis, and are generally converted into chromates, which are recognized by their color and the yellow precipitate of lead chromate formed when mixed with lead acetate.

Rhodium, Rh, 104.3, discovered by Wollaston in 1803, exists in platinum ores. It is hard and brittle, melting only at a high temperature, and then oxidizing. Specific gravity, 12.1. The atomicity of rhodium seems to resemble that of iron; it forms a monoxide and sesquioxide.

Iridium, Ir, 198, discovered by Tennant in 1803, is found with platinum and osmium. It is a hard, white, not very tough solid, which is difficult to fuse and to dissolve. It forms, like rhodium, compounds which recall those of iron. We have iridous chloride, $IrCl_2$, and iridic chloride, Ir_2Cl_6.

Osmium, Os, 199, discovered by Tennant in 1803, is associated with platinum. It resembles platinum in many properties, and has high lustre; specific gravity, 21.4; very difficult to fuse, and forms an amalgam with mercury. Five oxides are known, OsO, Os_2O_3, OsO_2, OsO_3, OsO_4. The monoxide and sesquioxide form an imperfect series of salts; the trioxide and tetroxide form salts with bases.

Molybdenum, Mo, 96, discovered by Bergman in 1781, occurs chiefly as disulphide and as lead molybdate, both rather rare minerals. It forms three oxides, MoO, MoO_2, MoO_3. The last, *molybdic anhydride*, forms salts called molybdates. Ammonium molybdate, Am_2MoO_4, is used as a test for phosphoric acid, with which it forms a highly insoluble yellow precipitate.

Uranium, U, 120, was discovered by Klaproth in 1789. It is rather rare, occurring as oxide and phosphate. It is a grayish solid, not oxidized by air or water, but dissolving in

acids. It forms two classes of compounds, uranous (tetrad) and uranic (hexad). Uranic oxide, UO_3 can act as an anhydride, forming uranates. Uranium compounds are generally yellow or yellowish-green, and are *fluorescent;* that is, shine brilliantly under the influence of rays of light to which the eye is ordinarily not sensitive. Glass colored yellowish with uranium oxide is much used in optical experiments.

The following elements are not yet definitely classified, but are probably related to aluminum:

Cerium, Ce, **Lanthanum,** La, and **Didymium,** Di, occur together in a few rare minerals. They form oxides which are analogous to alumina, and highly insoluble oxalates. Solutions of didymium salts when examined by the spectroscope show several *absorption bands;* that is, points at which the light is interrupted.

Indium, In, 113.4, exists in ores of zinc obtained at Freiberg, Saxony. It was discovered in 1863 by Reich and Richter. It gives when heated a violet-blue flame, which is a mixture of indigo and blue.

Gallium, Ga, 68.9, exists in small quantities in various zinc ores. It has only been obtained in small quantity. It was discovered by Lecoq de Boisbaudran in 1875. Gallium melts at 86° F. (30° C.), and will therefore melt by the heat of the hand.

16 *

ORGANIC CHEMISTRY.

Organic Bodies were at first understood to be those that exist only in living structures. The progress of chemistry made known many substances which could be produced by artificial means from the true organic bodies, and thus organic chemistry came to include not only the constituents of animals and plants, but the derivatives from them. A still further advance was made when some of these substances were produced directly from their elements without the intervention of life, thus showing that inorganic and organic bodies were not essentially different. Attempts have been made of late years to dispense with the separate consideration of organic bodies, and to include their description under the title " Carbon Compounds," on account of the almost invariable presence in them of that element. The method has not reached general acceptance, and to follow it at present would be of doubtful scientific advantage; and although the essential distinction between the two branches of the science has faded away, it is still convenient to describe under one section a great many bodies which *can* be made artificially, but which are generally derived from organized tissues. These so-called organic bodies are distinguished by a few peculiarities. The greater number of them contain carbon, hydrogen and oxygen. Some contain only carbon and hydrogen; others contain nitrogen. The extensive variety of compounds is due to differences in proportion and arrangement of atoms. The organized tissues of

plants and animals—muscle, brain, skin, seed, leaf, pollen, etc.
—are generally complex in character, and often contain other
elements, especially sulphur and phosphorus, in addition to
those mentioned above. Many organic bodies are subject to
change, and, as a rule, the more complex the substance, the
more easily is it decomposed.

Organic chemistry presents us with a large number of
radicles. The most important of these are compounds of
carbon and hydrogen. Their number is greatly increased
by the power which the carbon possesses (at least this is the
accepted theory) of combining with itself, forming duplicated
atoms, or, as they have been called, *carbon skeletons.* In this
manner we have

Carbon, tetrad.　　Dicarbon, hexad.　　Tricarbon, octad.

$$-\overset{|}{\underset{|}{C}}-\qquad -\overset{|}{\underset{|}{C}}-\overset{|}{\underset{|}{C}}-\qquad -\overset{|}{\underset{|}{C}}-\overset{|}{\underset{|}{C}}-\overset{|}{\underset{|}{C}}-$$
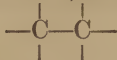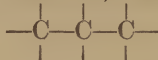
It will be noticed that the addition of each carbon atom adds
two degrees to the atomicity, to saturate which two atoms of
hydrogen will be required. It follows, therefore, that if we
add CH_2 to any radicle, we do not change its atomicity. In
CH_3, for instance, we have a monad radicle of which the
graphic formula may be given $-C\equiv H_3$; if to this we add
$=C=H_2$, we will get $H_2=C-C\equiv H_3$, in which one bond is
still unsatisfied. In accordance with this principle, we have
a number of series, the members of which differ by CH_2, and
possess the atomicity of the lowest member. Such a series,
differing by CH_2, is said to be HOMOLOGOUS.

It is convenient to arrange organic compounds according
to the radicles they are supposed to contain, although a great
many bodies cannot yet be referred to known radicles. These
latter will either be described apart or in connection with those
of known composition that they most nearly resemble.

The complexity of organic compounds is such that two
bodies may have the same composition, but have only par-

tial resemblance or even be entirely different in properties. This is called ISOMERISM, and is generally explained by supposing that the atoms are differently arranged in the different bodies. The annexed cuts will serve to represent the general principle of this explanation. Each block is made up of the same number of squares.

Sometimes the difference of properties is explained by the supposition that the *number* of atoms present in the one substance is greater than in the other, although the *proportion* between the different atoms is the same. Thus, the following compounds are known:

$$CNCl \qquad \text{gaseous.}$$
$$C_2N_2Cl_2 \qquad \text{liquid.}$$
$$C_3N_3Cl_3 \qquad \text{solid.}$$

In these it will be seen that the percentage composition is the same. Bodies that are related in this manner are called *polymers.*

Organic bodies, as they occur in nature, are generally mixtures of several distinct substances which may be very different in properties. Common fats, for instance, are mixtures of three distinct fats; common rosin contains at least two, and often three, substances.

The separation of these bodies from each other is called PROXIMATE ANALYSIS, and the bodies so separated are called PROXIMATE PRINCIPLES. The number of organic compounds has been greatly increased by the action of chemical agents upon these natural proximate principles. The chief means of producing such changes are briefly described:

(*a*) *Action of Heat.* Heat usually decomposes organic substances, leaving a residue of carbon and producing a mixture of bodies—solid, liquid and gaseous.

The general nature of this effect may be observed by placing some pieces of bituminous coal in a test-tube provided with a cork and glass jet, and heating to redness. The escaping gas is similar in composition to common illuminating gas, but burns with a rather smoky flame on account of the tar present.

(*b*) *Action of Oxygen.* Oxygen when acting at low temperature generally forms acids, either by direct union or by removing the hydrogen and taking its place. A mixture of potassium permanganate and caustic alkali is now much used as an oxidizing agent.

(*c*) *Action of Nitric Acid.* This sometimes produces a simple oxidation. Very often it removes hydrogen and adds NO_2. Nitrous acid sometimes removes H_3 and substitutes N.

(*d*) *Action of Chlorine.* Chlorine generally removes hydrogen, and takes its place, atom for atom. Bromine and iodine have the same action.

This substituting action of nitric acid and halogens gives rise to a large series of compounds, organic in general relations, but not existing in any living tissue. The substitution of NO_2 gives us a series of nitro-compounds, and takes place in the proportion of one molecule of NO_2 for each atom of H. Nascent hydrogen often removes the NO_2, and brings the body back to its original condition.

(*e*) *Action of Dehydrating Agents.* Dehydrating agents (bodies which have strong affinity for water) generally remove both hydrogen and oxygen from the organic compound in the proportion of two atoms of H to one of O (H_2O).

(*f*) *So-called Natural Changes.* These are FERMENTATION, PUTREFACTION and DECAY. Fermentation is a change produced in an organic body by the action of a decomposing nitrogenous body whereby it is reduced to a simpler form. Putrefaction is a change taking place, especially in nitrogenous bodies, merely under the influence of the ordinary conditions. Decay is a modified putrefaction in which oxidation occurs.

The processes of fermentation and putrefaction are attended by the development of microscopic living organisms, which have of late years been extensively studied, and are by many supposed to be the cause of the changes. Pasteur believes that the organisms found in fermenting liquids grow and multiply at the expense of the fermenting body, and thus bring about the change. Liebig thought that the cause of the fermentation was that the decomposition going on in the nitrogenous body present is communicated, by contact, to the molecules of the other substances. These views have been extensively discussed in connection with the now popular ideas that most fevers and contagious diseases are caused by living organisms and germs; but the question really belongs to physiology and biology, and not to chemistry.

A considerable number of substances—*e. g.* zinc chloride, carbolic acid, salicylic acid, kreasote, etc.—have the power to prevent fermentation or putrefaction, and are called ANTISEPTICS.

Analysis of Organic Bodies.—The best general test for an organic substance is the action of heat, which usually causes decomposition, with evolution of smoky, strong-smelling va-

pors, and leaves a residue of carbon which can be burned off by heating strongly in the air.

This test can be easily applied by heating the substance on platinum-foil or a piece of porcelain, as shown in the cut.

Sulphuric acid also produces a characteristic blackening, due to liberation of carbon.

The presence of nitrogen is usually indicated by a very disagreeable odor on heating; more surely by heating the body with an alkali, by which ammonia is formed. The accurate analysis of organic bodies is performed by burning them completely in a current of oxygen, and collecting and weighing the carbon dioxide, water and nitrogen which are thus produced. This process gives the *percentage* or *ultimate* composition.

A formula which expresses merely the composition of the body, without exhibiting the known or supposed arrangement of the atoms, is called an EMPIRICAL formula; a RATIONAL formula is one which shows the arrangement of the atoms. Butyric acid and acetic ether are represented by the same empirical formula, $C_4H_8O_2$, but their rational formulæ show their complete difference:

$C_2H_5C_2H_3O_2$ Ethyl acetate (acetic ether).
$HC_4H_7O_2$ Hydrogen butyrate (butyric acid).
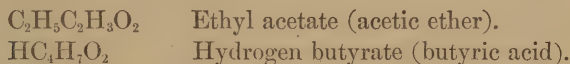
PROXIMATE organic analysis, or the separation of the bodies existing in a mixture, is a matter entirely dependent on the nature of the substances present; it is sometimes very easy, and in other cases so difficult as to be practically impossible. Solids are generally separated by differences of solubility in water, alcohol, ether and other solvents; liquids are separated by *fractional distillation*, in which advantage is taken of different boiling-points. Gases are separated by different absorbents. The principle of these methods may be illustrated by making a mixture of starch, sugar, rosin and some volatile oil. By gently heating the mixture the oil may be distilled off; cold water will take up the sugar, and hot water the starch; alcohol will then dissolve the rosin.

NOMENCLATURE OF ORGANIC BODIES.

Great difficulty has been found in the nomenclature of organic bodies, and many have been given arbitrary names derived from fanciful relations or peculiarities. A large number of proximate principles are named from the articles from which they are extracted. The sugar in milk is called *lactose*, from *lac*, " milk ;" the active principle of tobacco is called *nicotina*, from *nicotiana*, the botanical name of the tobacco, etc.

A large number of organic bodies may be regarded as made up of positive and negative radicles, in the same manner as inorganic bodies are made up of positive and negative elements ; and by assigning names to these radicles we form the names of the compounds into which they enter, just as we form the names of inorganic compounds. When the compound contains an acid radicle, the syllable " *ate* " is used.

In all cases in which a body has been formed by substitution of one element or molecule for another, we may easily indicate the fact by attaching the name of the substituting body, and using the prefixes MONO, DI, TER, etc. to show the amount. To indicate the substitution of the molecule NO_2, the word *nitro* is used ; to indicate the substitution of N for H_3, the word *azo* is used.

The application of these principles will be shown in the following pages.

Classification of Organic Bodies. In the following pages the substances described will, as far as possible, be arranged according to the radicles supposed to be present, beginning, as in inorganic chemistry, with the monads, and proceeding to those of higher atomicity.

ORGANIC BODIES NOT CONTAINING NITROGEN.

Hydrocarbons and Derivatives.—Bodies consisting only of hydrogen and carbon are called *hydrocarbons.* Many of these are produced by the destructive distillation of organic substances, especially coal and wood. They are arranged in a number of homologous series, of which only the more important will be given. The fusing- and boiling-points gradually rise as the number of carbon atoms increases, so that if the lower members are gases, those in the middle of the series will be liquids, the higher members will be solids. In each series the atomicity is determined by that of the first member, because, as explained above, the addition of CH_2 does not affect the saturation.

The series begins with CH_4, which, being saturated, has no disposition to combine, and therefore is not truly a radicle. By successive subtractions of hydrogen we get radicles of increasing atomicity. The following table shows this fact, and also shows the system of nomenclature:

(Saturated.)	(Monad.)	(Dyad.)	(Triad.)	
Methane	methyl	methene	methenyl	
CH_4	CH_3	CH_2	CH	
Ethane	ethyl	ethene	ethenyl	ethine
C_2H_6	C_2H_5	C_2H_4	C_2H_3	C_2H_2
Propane	propyl	propene	propenyl	propine
C_3H_8	C_3H_7	C_3H_6	C_3H_5	C_3H_4
Butane	butyl	butene	butenyl	butine
C_4H_{10}	C_4H_9	C_4H_8	C_4H_7	C_4H_6
Quintane	quintyl	quintene	quintenyl	quintine
C_5H_{12}	C_5H_{11}	C_5H_{10}	C_5H_9	C_5H_{10}

The gradual diminution of hydrogen is indicated by the change of vowel in regular order; those containing uneven numbers of hydrogen have, in addition, the syllable "*yl*,"

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The atoms of carbon are (except in the first four, which have arbitrary names) indicated by the syllable taken from Greek or Latin numerals.

Series differing by H_2, but containing the same number of carbon atoms, are called ISOLOGOUS. Thus C_3H_8, C_3H_6, C_3H_4, are isologous.

1st Series, METHANES, *often called* PARAFFINS. These begin with CH_4, and being saturated bodies they are comparatively indifferent to chemical reagents.

Marsh Gas, or *Methane*, CH_4, is produced by distillation of wood and coal or by decay of vegetable matter. It is, next to hydrogen, the lightest body known. It exists in coal gas.

Common Paraffin exists in petroleum and in coal-tar. It is a mixture of several of the higher members of the series. It is a white, waxy solid, easily fusible, little acted on by acids or alkalies, and used for a protecting coating in chemical apparatus. Cosmoline and its imitations are also in part soft paraffins.

2d Series, METHYLS. These begin with methyl, CH_3, and are a series of monad radicles, forming the monatomic alcohols.

Some of the important members of the series are here given :

CH_3	Methyl.	C_3H_7	Propyl.
C_2H_5	Ethyl.	C_4H_9	Butyl.
C_5H_{11}	Amyl, or Quintyl.		

These form compounds analogous in structure, but not in appearance or general properties. The compounds are very important. We have—

1. Normal oxides, called SIMPLE ETHERS :

$(CH_3)_2O$ Methyl ether, analogous to Na_2O, sodium oxide.
$(C_2H_5)_2O$ Ethyl ether   "   "   "   "

2. Compounds with halogens, sometimes called ETHERS :

$(CH_3)Cl$ Methyl chloride, analogous to $NaCl$, sodium chloride.
$(C_5H_{11})Cl$ Amyl chloride,   "   "   "   "

3. Compounds analogous to salts, called COMPOUND ETHERS:

$(CH_3)_2SO_4$ Methyl sulphate, analogous to Na_2SO_4, sodium sulphate.
$(C_5H_{11})NO_3$ Amyl nitrate,   " $NaNO_3$,   "   nitrate.

4. Compounds analogous to the acid salts, called VINIC ACIDS:

$(C_2H_5)HSO_4$ Sulph-ethylic acid, analogous to $KHSO_4$.
$(C_5H_{11})HSO_4$ Sulph-amylic acid,   "   " .

5. Compounds analogous to the hydrates, called ALCOHOLS:

$(C_2H_5)HO$ Ethyl alcohol, analogous to KHO.
$(C_5H_{11})HO$ Amyl alcohol,   "   "

6. Compounds containing two different radicles, analogous to the mixed salts and called MIXED ETHERS:

$(CH_3)(C_2H_5)O$ Methyl-ethyl ether.

Derivatives from the Second Series of Radicles.—

Some of the derivatives from this series have been known in more or less impure form for centuries; one of them, common alcohol, being the active agent in intoxicating liquors, the use of which in one form or other dates from remote antiquity. The term *alcohol* is now extended to all the hydrates of positive radicles.

Methyl Alcohol, Wood Spirit, $(CH_3)HO$, *Methyl Hydrate,* is usually made by distilling wood. The crude material is difficult to purify. Pure methyl alcohol is colorless and of pleasant odor. It boils at 152° F. (66.5° C.), and its effects on the animal system appear to be less severe and more transient than those of common alcohol. The methylated spirit of English chemists is a mixture of 90 parts common alcohol with 10 parts methyl alcohol.

Ethyl Alcohol, Common Alcohol, Spirit of Wine, $(C_2H_5)HO$, *Ethyl Hydrate,* is produced in the vinous fermentation of sugar; alcohol and carbonic anhydride being chiefly formed: it can also be prepared artificially. On the large scale the sprouted grain called malt is generally used. The general nature of fermentation will be explained in connection with the sugars. The fermented spirit, having a lower boiling-point than water, is concentrated by distillation, but the strongest

spirit thus prepared contains 10 per cent. of water. To withdraw all the water, it is necessary to distil with quicklime, by which *absolute alcohol* is formed. This is very inflammable, greedily absorbs moisture, and mixes with water in all proportions.

Proof spirit contains 50.8 parts by weight of alcohol to 49.2 of water, and has a specific gravity of 0.920. Commercial alcohol is a colorless, volatile liquid, of which the properties, effects and uses are well known. The strongest spirit ordinarily furnished is about 95 per cent., and boils at about 180° F. (81° C.). Alcohol is contained in wine, beer and spirits; essential oils, sugar or extracts are mixed with it as flavoring agents. Whiskey, brandy and other spirits contain from 40 to 50 per cent. of alcohol; wines, from 17 (port and madeira) to 7 or 8 (hock and light clarets) per cent.; porter and strong ale contain from 6 to 8 per cent.; lager beer, about 4 per cent. The effervescence of fermented liquids is due to the carbon dioxide which is produced with the alcohol, thus:

$$\text{Glucose,} \qquad\qquad \text{Alcohol,}$$
$$C_6H_{12}O_6, \text{ breaks up into } 2C_2H_6O + 2CO_2.$$
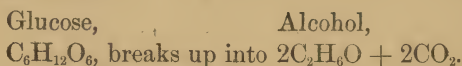
The carbon dioxide is retained by bottling the liquid before the fermentation is over.

Amyl Alcohol, Fusel Oil, $(C_5H_{11})HO$, *Amyl hydrate,* is a by-product in fermentation, and is found in raw spirits and new liquors. When pure it is a colorless, oily liquid, with a peculiar odor, a hot and acrid taste, and decidedly poisonous in its action.

The alcohols derived from the higher radicles are mostly wax-like.

Ether, Ethyl Oxide, $(C_2H_5)_2O$, is made by the action of dehydrating agents, especially sulphuric acid, upon alcohol. It appears that acid ethylsulphate is first formed and then decomposed.

$$\text{Alcohol,} \qquad\qquad \text{Acid ethylsulphate,}$$
$$(C_2H_5)HO + H_2SO_4 = (C_2H_5)HSO_4 + H_2O.$$

Another molecule of alcohol is then acted upon, thus:

Ether,

$$(C_2H_5)HO + (C_2H_5)HSO_4 = H_2SO_4 + (C_2H_5)_2O.$$

Ether is a colorless, very volatile liquid of distinct odor, boiling at 95° F. (35° C.). Specific gravity, 0.713. Its vapor is inflammable and very heavy. It is a solvent for fats, fixed and volatile oils, resins and many other proximate principles. Methyl and amyl ethers are also known.

Compound Ethers. The alcohol radicles replace the hydrogen of acids and form bodies called compound ethers. It will be sufficient to enumerate a few of the most important. Many of these compound ethers are fragrant, and are used as flavoring materials or as substitutes for fruit-essences.

Ethyl bromide, $(C_2H_5)Br$, is used in medicine. Amyl acetate, butyrate and valerate are used as flavors. Amyl nitrite is used in medicine. Ethyl nitrate exists in sweet spirit of nitre.

ALDEHYDES AND ACIDS. When alcohols are oxidized by a limited amount of oxygen, two atoms of hydrogen are removed and no oxygen is added. When oxidized in a free supply of oxygen, an atom of oxygen takes the place of the removed hydrogen. The bodies produced in the first case are *aldehydes;* in the second, *acids.* In this way we have

$$(CH_3)HO + O = CH_2O \qquad \text{Methyl aldehyde.}$$
$$(CH_3)HO + O_2 = CH_2O_2 \qquad \text{Formic acid.}$$
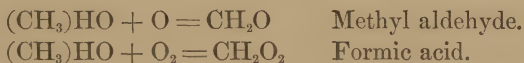
Thus each alcohol may be made to yield an aldehyde and an acid. The acids are very important, and are often called *fat acids*, many of them existing in common fats. They may be produced by the substitution of one atom of oxygen for two atoms of hydrogen in the alcohol. The removed hydrogen is converted into water by another atom of oxygen. Thus,

$$\text{Methyl Alcohol,} \qquad \text{Formic Acid,}$$
$$(CH_3)HO + O_2 \quad = \quad CH_2O_2 + H_2O.$$

The acids formed from this series of radicles are *monobasic;*

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that is, only one atom of hydrogen is capable of being replaced by a metal. They are therefore usually written with a single atom of hydrogen separate from the rest; formic acid, for instance, is written $HCHO_2$. The following are some of these acids and their corresponding alcohols:

Alcohol.				Acid.	
Methyl alcohol,		$(CH_3)HO$,	yields	$HCHO_2$,	Formic acid.
Ethyl	"	$(C_2H_5)HO$,	"	$HC_2H_3O_2$,	Acetic acid.
Propyl	"	$(C_3H_7)HO$,	"	$HC_3H_5O_2$,	Propionic acid.
Butyl	"	$(C_4H_9)HO$,	"	$HC_4H_7O_2$,	Butyric acid.
Amyl	"	$(C_5H_{11})HO$,	"	$HC_5H_9O_2$,	Valeric acid.

Only a few of this series of acids will need description.

FORMIC ACID, $HCHO_2$, exists as a secretion in some stinging animals and plants. It is a corrosive volatile liquid and acts as a reducing agent.

ACETIC ACID, $HC_2H_3O_2$. This, in the dilute state, constitutes vinegar, which contains about 5 per cent. of the acid, and is usually made by oxidizing dilute alcohol in the presence of a ferment. Acetic acid is also produced in the distillation of wood. When pure, it is a colorless, corrosive liquid, solidifying at 62.6° F. (17° C.), and boiling at 246° F. (119° C.). This is *glacial* acetic acid. The more dilute forms are less active, and in vinegar its effects are quite mild.

Acetates. Many of these are important. *Potassium* and *sodium acetates* are very deliquescent.

Lead acetate, $Pb(C_2H_3O_2)_2$, *Sugar of Lead,* made by dissolving lead oxide in acetic acid, forms white crystals soluble in water. The solution is capable of taking up more lead oxide, and forming an oxyacetate known as Goulard's extract.

Cupric oxyacetate is a compound of copper acetate with copper hydrate, and is known as *verdigris.*

VALERIC ACID. This is obtained by the action of oxidizing agents on amylic alcohol. It also exists in some plants.

It is an oily liquid of disagreeable odor. Several of its salts are in extensive use in medicine.

The higher members of this monobasic series are generally insoluble in water and oily or fatty in appearance.

Third Series, OLEFINS. These begin with C_2H_4, olefiant gas, ethylene, which is contained in coal gas, and can be made pure by the action of dehydrating agents on common alcohol. The olefins are dyads. The hydrates constitute the *diatomic alcohols* or *glycols*. These are not important. *Ethene glycol*, $(C_2H_4)H_2O_2$, is the best known. It is a thick liquid, sweet and colorless.

ACIDS FROM GLYCOL. Two series of these exist—one derived by the replacement of *two* atoms of hydrogen by one atom of oxygen, and the other by the replacement of *four* atoms of hydrogen by two atoms of oxygen. The first is the *lactic acid series;* the second, the *oxalic acid series.*

The following will show the relation between the two groups:

Glycol.	First oxidation.	Second oxidation.
$C_2H_6O_2$	$C_2H_4O_3$, Glycolic	$C_2H_2O_4$, Oxalic.
$C_3H_8O_2$	$C_3H_6O_3$, Lactic	$C_3H_4O_4$, Malonic.
$C_4H_{10}O_2$	$C_4H_8O_3$, Oxybutyric	$C_4H_6O_4$, Succinic.

Each series is homologous, and carbonic acid, CH_2O_3, may be regarded as the first member of the first series.

Only a few of these acids need description.

Lactic acid, $HC_3H_5O_3$, is contained in sour milk, and is formed from sugar during the so-called lactic fermentation. It may be prepared artificially. It is a syrupy liquid which forms soluble salts.

Oxalic acid, $H_2C_2O_4$, exists ready formed in the juices of some plants, as sour-sorrel and garden rhubarb. It is made artificially by the action of nitric acid on sugar or of caustic alkali on sawdust. It is sometimes a result of disease in the animal system. It occurs in colorless crystals which contain

$2H_2O$, and is soluble in water, very sour and highly poisonous. Many of its salts are insoluble in water, especially calcium oxalate. Its decomposition with sulphuric acid is a convenient source of carbon monoxide (page 139). With the monads two oxalates are formed. *Acid potassium oxalate,* KHC_2O_4, occurs in the leaves of plants, and is called salt of sorrel. *Calcium oxalate,* CaC_2O_4, is occasionally formed in the kidneys. Ammonium oxalate, $Am_2C_2O_4$, is used as a test.

Succinic acid is produced from amber, a fossil resin, and exists in a few plants. It can also be made by the action of nitric acid upon most fats.

Fourth Series, METHENYL SERIES. These are triad radicles, beginning with CH, methenyl, which forms $(CH)Cl_3$, analogous to $BiCl_3$.

The series may be regarded as forming the *triatomic alcohols* or *glycerins.*

$(C_3H_5)H_3O_3$, glycerin, is the best known of the series. It is obtained by the decomposition of fats. It does not oxidize or evaporate in the air, and has a solvent action next to that of water. Heated with strong nitric acid, it forms nitroglycerin, $C_3H_5(NO_2)_3O_3$ (strictly trinitro-glycerin), which is a violent explosive, and when mixed with sand or other material constitutes dynamite.

Chloroform, $CHCl_3$, can be obtained by the action of chlorine on marsh gas, but the commercial article is made by distilling alcohol with bleaching-powder. The reaction is complicated, and the product is at first quite impure. When pure it is a colorless, fragrant liquid, very volatile, not easy to burn, insoluble in water and much heavier than that liquid. It boils at 142° F. (61° C.). It has high solvent powers and is a valuable anæsthetic.

Fats and Fixed Oils. These are compound ethers, mostly derived from *propenyl,* (C_3H_5), and therefore, like bismuth, requiring three molecules of a monobasic acid to

form a normal salt. As found in the tissues of plants and animals, they are mixtures of several distinct bodies which can be separated by difference of melting-point, solubility, etc. The oils are merely fats with a low melting-point. They are usually divided into two classes: *drying oils*, which absorb oxygen from the air, and become hard and resinous, such as linseed and poppy oil; *non-drying oils*, which remain fluid, as castor and sperm oil. Many fats and oils undergo partial decomposition in the air, producing a free acid; this is called *rancidity*.

Each distinct fat is usually called by a name derived from the acid that forms it. Glyceryl oleate is called *olein;* glyceryl stearate is called *stearin.* When caustic alkali is added to a fat, decomposition takes place; *soap* and *glycerin* are produced. The reaction is analogous to that which occurs when bismuth nitrate is acted upon by an alkali.

$$Bi(NO_3)_3 + 3KHO = BiH_3O_3 + 3KNO_3.$$

Stearin,   Potass. stearate,   Glycerin,
$$(C_3H_5)(C_{18}H_{35}O_2)_3 + 3KHO = 3K(C_{18}H_{35}O_2) + (C_3H_5)H_3O_3.$$
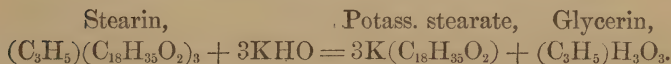
Soaps produced by potassa are usually soft; those from soda, hard; those made from other oxides are mostly insoluble in water. This latter fact explains the curdling action of hard water. The calcium and magnesium salts produce insoluble soaps. When soluble soaps are treated with cold water they decompose into acid salt, which precipitates and makes the soapsuds, and an oxy-salt, which dissolves and gives the cleansing action.

A decomposition of the fats may be produced by the action of superheated steam. In this case glycerin and free acids are formed, and can be distilled off. The process is used on the large scale.

Stearin,   Glycerin,   Stearic acid,
$$(C_3H_5)(C_{18}H_{35}O_2)_3 + 3H_2O = (C_3H_5)H_3O_3 + 3H(C_{18}H_{35}O_2).$$
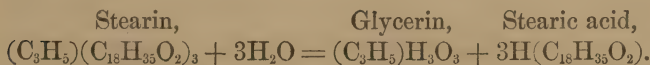
The acids of the fats may also be obtained by adding a strong acid to ordinary soaps.

Of the higher series of hydrocarbons, the following are the most important:

BENZENES. These begin with C_6H_6, benzene, benzole or coal-tar naphtha, which is the important member.

Benzole or *benzene* is obtained by the fractional distillation of coal-tar and by the distillation of benzoic acid with lime. It is a colorless, volatile liquid, which solidifies at a little above the temperature of melting ice. Its solvent powers are very high, and it is largely used for such purpose. Benzene must not be confounded with the mixture of hydrocarbons obtained by distilling coal oil and known as benzine.

Nitrobenzole, oil of myrbane, $C_6H_5(NO_2)$, is obtained by the action of strong nitric acid upon benzene:

$$C_6H_6 + HNO_3 = C_6H_5(NO_2) + H_2O.$$

The product is a yellow, oily liquid smelling like bitter almonds, insoluble in water and poisonous. It is important, because by the action of nascent hydrogen it yields aniline, from which many brilliant colors are made.

$$\text{Nitrobenzole,} \qquad \text{Aniline,}$$
$$C_6H_5(NO_2) + H_6 = (C_6H_5)H_2N + 2H_2O.$$

NAPHTHALENES.—*Naphthalene*, $C_{10}H_8$, the only important member of the group, occurs in coal-tar. It is a white, crystalline solid, melting at 176° F. (80° C.), and slightly soluble in water.

ANTHRACENES.—*Anthracene*, $C_{14}H_{10}$, is obtained from coal-tar. It is especially valuable as a source of alizarine, the coloring-matter of madder. When pure it is a colorless solid, subliming at 212° F. (100° C.), insoluble in water.

Anthraquinone, $C_{14}H_8O_2$, is obtained by the action of a powerful oxidizing agent, as chromic acid, upon anthracene:

$$C_{14}H_{10} + O_3 = C_{14}H_8O_2 + H_2O.$$
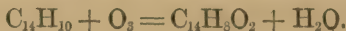
Anthraquinone is used as a source of artificial madder color. The process is, (*a*) the anthraquinone is heated with bromine:

Dibromanthraquinone,

$$C_{14}H_8O_2 + Br_4 = C_{14}H_6Br_2O_2 + 2HBr.$$
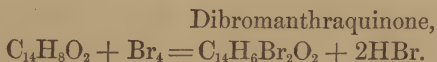
(*b*) Dibromanthraquinone heated with potash gives potassium alizarate, from which the alizarine may be obtained by sulphuric acid.

Potass. alizarate,

$$C_{14}H_8Br_2O_2 + 4KHO = C_{14}H_6K_2O_4 + 2KBr + 2H_2O.$$

TERPENES. These are principally natural products constituting the *volatile or essential oils.* C_5H_6, called quintone, is the lowest member of the series known.

Oil of Turpentine, $C_{10}H_{16}$, is obtained from turpentine, an exudation from pine trees, and consisting of resin and volatile oil. On being distilled, the volatile oil is collected in the receiver; the resin remaining constitutes common rosin. Oil or spirits of turpentine is a thin, colorless liquid of peculiar odor. It is lighter than water, boils at 320° F. (160° C.) and is a valuable solvent. It is partially oxidized in the air. Many fragrant oils obtained from plants have the same composition, and are obtained by distillation. Some of these oils are—lemon, bergamot, coriander, hop, juniper and valerian. They are called *essential oils*—are mostly lighter than water and freely soluble in alcohol and ether.

Plants furnish us with a number of oxidized terpenes, among which are the *camphors* and *resins.*

Common Camphor, $C_{10}H_{16}O$, obtained from the camphor laurel, is a white crystalline solid, volatile at ordinary temperatures. It is slightly soluble in water and freely in alcohol and ether.

Resins include a large group, of which many are true acids, and form salts constituting resin soaps.

Common Rosin is the residue from the preparation of oil of turpentine. It is a mixture of two acids.

Lac, Copal and *Mastich* are familiar members of the group. As a class the resins are easily fusible, but not volatile; insoluble in water, but soluble in alcohol, which solution constitutes varnish. Gum-resins are simply mixtures of resin and gum; oleoresins, mixtures of resin and volatile oil; balsams contain benzoic or cinnamic acid.

Caoutchouc and *Gutta Percha* are terpenes found in the juices of plants. They are insoluble in water, but in the plants are usually in suspension, very finely divided, so as to make a milky liquid called an emulsion. Caoutchouc is elastic; gutta percha is not. Both are capable of combining with sulphur to form peculiar and valuable compounds. The process is called vulcanizing.

SUGARS AND STARCHES.

The sugars form an important group, the exact relations of which are not well understood, but they are generally regarded as alcohols or aldehydes derived from complex radicles. The group includes sugars proper, also gums, starches and woodfibre. A remarkable similarity of composition and convertibility into one another by simple means is to be noted. The most important point in regard to their composition is that they contain oxygen and hydrogen in the proportion to form water. They are divided into three classes:

1. SUCROSES (sugars proper).	2. GLUCOSES (grape sugars).	3. AMYLOSES (starch and woody fibre).
$C_{12}H_{22}O_{11}$	$C_6H_{12}O_6$	$C_6H_{10}O_5$
Sucrose (Cane Sugar).	Dextrose (Grape Sugar).	Starch.
Lactose (Milk Sugar).	Lævulose (Fruit Sugar).	Dextrin.
		Inulin.
		Gum.
		Cellulose.

Sucrose, or *Cane Sugar,* $C_{12}H_{22}O_{11}$, exists in certain plants,

especially sugar-cane and beet-root. It is extracted by press-
ure; the liquid is then boiled down carefully, the raw prod-
uct decolorized by animal charcoal and finally crystallized.
It is soluble in about twice its weight of cold water. When
heated to about 420° F. (216° C.) a caramel is formed.

Lactose, or *Milk Sugar*, is found only in milk. It is con-
verted by dilute acids into a peculiar variety of glucose, and
in the presence of cheese undergoes lactic fermentation.

Glucose, or *Grape Sugar*, $C_6H_{12}O_6$, is found in manna and
honey and many kinds of fruits. It is a normal constituent
of the blood, and is excreted in considerable amount in the
disease called *diabetes mellitus*.

Glucose presents two modifications, dextrose and lævulose,
distinguished by their action on light. Dextrose may be
obtained by boiling starch with dilute sulphuric acid, add-
ing chalk and evaporating the liquid. It is soluble in dilute
alcohol, but is not so sweet as sucrose. When a solution of
cane sugar is boiled with dilute acids, a mixture of dextrose
and lævulose is formed, called *inverted sugar*. Lævulose does
not crystallize, and is sweeter than dextrose.

Starch, $C_6H_{10}O_4$, occurs in many plants. It is a white pow-
der, which is made up of granules of various sizes having a
definite organized structure. (The annexed
cut shows a magnified view of the cells of the
potato, with the starch granules in position.)
These granules are not soluble in cold water,
ether or alcohol, but if heated with water to
about 160° F. (72° C.) they swell and break
up, yielding a thick mass termed starch paste.
Upon boiling this mass with more water, the
particles are reduced to so fine a state of division that they
will pass through a filter, and when the boiling is continued
for some time the solution becomes clear and the starch solu-
ble. The test for starch is the formation of a deep blue color
with free iodine. Starch exists in the seeds of grasses, asso-

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ciated with an albuminous substance, *diastase*, which has the power to transform the starch into glucose. When the seed germinates this transformation begins, and if the germination be interrupted before the sugar begins to undergo further change, we have *malt*, which is simply sprouted grain, especially barley. When malt is steeped in water and yeast added, the fermentation of the sugar begins. Dilute sulphuric acid acts like diastase.

Dextrin, $C_6H_{10}O_5$. This substance is also known as British gum, and may be obtained by heating starch to about 320° F. (160° C.). The change is much more speedily effected by the addition of a little hydrochloric or nitric acid. Dextrin, together with dextrose, is formed when malt extract acts upon starch. It is insoluble in alcohol, but very soluble in water, and is used as a mucilage. It is converted into glucose by heating with dilute acids.

Gum Arabic is a natural exudation from many species of *Acacia*. It consists chiefly of arabic acid, $C_{12}H_{20}O_{10}$, united with calcium and potassium.

Cellulose, $C_6H_{10}O_5$, is the colorless material of woody fibre. It is obtained nearly pure by boiling cotton with alkali. Cellulose is a white substance, which dissolves in an ammoniacal solution of cupric oxide, but is insoluble in water, ether or alcohol. Strong sulphuric acid converts it either into a soluble substance like dextrin, or into an insoluble substance, giving a blue color with iodine. By dipping sheets of paper into strong sulphuric acid *parchment paper* is obtained.

Gun-Cotton. When cotton is put into a mixture of equal volumes of strong nitric and sulphuric acids, no apparent change occurs, but after drying it is found to be exceedingly inflammable. A substitution product is here formed, termed *trinitro-cellulose*, in which NO_2 replaces hydrogen, $C_6H_7(NO_2)_3O_5$.

Collodion is formed by dissolving certain kinds of gun-

cotton in a mixture of ether and alcohol. It is much employed in photography and surgery.

Mannite, $C_6H_8(HO)_6$ or $C_6H_{11}O_6$, is a sugar-like substance contained in manna, an exudation from many species of ash. When acted upon by nitric acid it forms a compound termed nitro-mannite, or mannite in which six atoms of hydrogen are replaced by NO_2.

The sugars and starches are reducing agents, but not very energetic. The action is generally increased by the presence of strong alkalies. The tests for them are mostly dependent upon their reducing action upon the salts of copper, silver, bismuth and mercury. Glucose is especially active.

Moore's test depends on the fact that a solution of sugar becomes darker on being boiled with caustic alkali.

Trommer's test is performed by adding to the suspected solution a few drops of copper sulphate and a considerable amount of caustic soda or potassa. If sugar be present, the application of heat will cause the precipitation of orange-red cuprous oxide, Cu_2O.

Boettger's test is performed in the same way, substituting bismuth oxynitrate for copper sulphate. A black precipitate of free bismuth is formed. The most accurate test is by fermentation and recognition of the alcohol.

Glucosides. These occur in many plants, and on decomposition give rise to a glucose, together with other substances. Only a few need be mentioned.

Amygdalin, $C_{20}H_{27}NO_{11} + 3H_2O$, is found in bitter almonds. It forms white soluble crystals. By the action of water (H_2O) upon the amygdalin and the albuminous substance termed *synaptase* or *emulsin*, contained in the bruised bitter almond, a species of fermentation is set up, in which hydride of benzoyl (C_7H_6O), hydrocyanic acid (HCN), and glucose ($C_6H_{12}O_6$), are formed.

Salicin, $C_{13}H_{18}O_7$, is found in the pith of the poplar and willow. It crystallizes in white brilliant needles. It is soluble in water and alcohol, the solution having a strong bitter taste. It yields saligenin and glucose.

Tannins or *Tannic Acids* are the astringent principles of plants. Different forms exist in different plants; most of them yield, on treatment with acid, glucose and gallic acid. They form with gelatin an insoluble precipitate but little liable to decomposition, and with ferric salts a dark precipitate which remains for a long time suspended in water. Upon the first of these reactions depends the process of manufacturing leather, and upon the second the formation of common inks.

The source of common tannin is gall-nuts, excrescences formed on various plants by the punctures of an insect. Tannin is uncrystallizable, insoluble in pure ether, but dissolved by water and alcohol. This form of tannin is apparently not a glucoside. It has the composition $C_{14}H_{10}O_9$.

FERMENTATION.

This is a change which organic bodies, especially the sugars and starches, suffer under the influence of complicated substances termed ferments, giving products differing according to the nature of the fermented body and of the ferment.

The circumstances necessary for the action appear to be— (*a*) proper food, especially the ammoniacal salts and alkaline phosphates; (*b*) a temperature of from 60° to 100° F. (15° to 40° C.), since at other temperatures the vitality of the ferment is destroyed.

Mention has been made on a previous page of the various theories as to the nature of fermentation.

A so-called spontaneous fermentation takes place in wine,

beer, milk and other liquids. Those who accept Pasteur's view ascribe the action to the sporules or seeds of living bodies floating in the air. These, dropping into the liquid, propagate themselves, and during the act of growing evolve the products of fermentation. If the liquid be left in contact only with air which has been passed through à red-hot platinum tube, or if the air be filtered by passing it through cotton, fermentable liquids may be preserved for any length of time without undergoing change.

The principal forms of fermentation are five:

1. The vinous, producing chiefly alcohol, C_2H_6O, and carbonic anhydride, CO_2.
2. The acetous, " " acetic acid, $C_2H_4O_2$.
3. The lactic, " " lactic acid, $C_3H_6O_3$.
4. The butyric, " " butyric acid, $C_4H_8O_2$.
5. The mucous, " " gum and mannite.

ORGANIC ACIDS NOT REFERABLE TO ANY OF THE ABOVE SERIES.

A great many organic acids are known, and referred to different incomplete series that cannot be here enumerated. It will be sufficient to mention some important acids and indicate their sources and properties:

Malic acid, $H_2C_4H_4O_5$, occurs in the juices of many plants, as apples, pears, mountain-ash berries and tobacco-leaves. It may be made artificially. It is crystalline, sour, soluble in water and alcohol. Its salts have very little importance.

Tartaric acid, $H_2C_4H_4O_6$. This is found in many plants, but especially in grapes, where it exists as acid potassium tartrate, $KHC_4H_4O_6$. This salt is somewhat soluble in water, but scarcely soluble in dilute alcohol, and hence in the manufacture of wine, as the fermentation advances, the quantity

18 *

of alcohol increases, and the acid potassium tartrate deposits as a red mass called *argols*. This, being dissolved in hot water and crystallized, gives *cream of tartar* (called very wrongly, by some writers, cremor tartar).

Tartaric acid is a crystalline body, soluble in water and very sour. Its solution in water develops a fungous growth and decomposes. Several varieties of the acid are known.

Acid potassium tartrate is a white crystalline body, very sour and not very soluble in cold water. It is used in effervescing powders.

Potassium tartrate, $K_2C_4H_4O_6$, is called soluble tartar.

Sodio-potassium tartrate, $NaKC_4H_4O_6$, is known as *Rochelle salt*.

Tartar emetic is described on page 131.

Citric acid, $H_3C_6H_5O_7$, is the acid of lemons and oranges, and is also found in some other fruits. It is a crystalline body, very sour and easily soluble in water.

Oleic acid, $HC_{18}H_{33}O_2$, exists in most natural fats and non-drying oils. It is solid at 57° F. (14° C.). Above this temperature it is a clear liquid, lighter than water and insoluble in it, but soluble in alcohol and ether. Crude oleic acid, made by the decomposition of fats by steam, as mentioned on page 200, is used in soap-making under the name of *red oil*.

Benzoic acid, $HC_7H_5O_2$, is found in many resins, but chiefly in benzoin, by heating which the acid sublimes in white, pearly plates. Most of the salts formed by benzoic acid are soluble.

Salicylic acid, $HC_7H_5O_3$, is formed by the oxidation of salicin, the bitter principle of the willow. It may also be made by acting upon sodium phenylate with carbon dioxide. Salicylic acid forms white crystals not very soluble in cold water, but rather soluble in hot water and in alcohol. It is now used largely as a medicine and antiseptic.

The acid also occurs in the oil of wintergreen, methyl salicylate, $(CH_3)C_7H_5O_3$.

Carbolic acid, HC_6H_5O, is hardly a true acid, but is so called from a power of taking up bases and forming bodies like salts. It is preferably called *phenol* or *phenylic alcohol*. It exists in coal-tar, and can also be made by several processes. It forms colorless crystals; is very deliquescent and soluble in water. It melts at 93° F. (34° C.), and boils at 370° F. (187° C.). It has a peculiar odor, much like that of kreasote, which is a somewhat similar body obtained from wood-tar. Phenol, salicylic acid and benzoic acid are used as antiseptics, and possess the power to retard the development of many forms of microscopic life.

Picric acid, or *Tri-nitro phenol*, $C_6H_3(NO_2)_3O$, is a yellow crystalline substance, very soluble in water, and is formed when nitric acid acts upon phenol. It may also be obtained when nitric acid acts upon many other substances, as the skin, for instance. In the arts it is employed as a dye for silk and wool. It is monobasic, and some of its salts are explosive.

Gallic acid, $HC_7H_5O_5$, is obtained from tannin. It is used as an astringent. By heat it yields *pyrogallin*, $C_6H_6O_3$, sometimes called pyrogallic acid, which has the property of absorbing oxygen rapidly when mixed with a base, and is used as a test for oxygen and as a reducing agent in photography.

ORGANIC BODIES CONTAINING NITROGEN.

These may be roughly divided into three classes:

1. Those derived from or containing cyanogen, CN.

2. Those containing NO_2 in substitution for H, and called *nitro-compounds*. Such of these as are of importance are

described in connection with the substances from which they are obtained.

3. Those derived from ammonia, and called *amines* or *amides.*

CYANOGEN AND DERIVATIVES.

Cyanogen, CN, is produced when organic matter containing nitrogen, such as leather scraps, is heated with an alkali, especially in the presence of iron. Cyanogen is a gas, but is of no importance in the free state. It is often represented by the symbol Cy (see page 147).

Hydrogen cyanide, HCN, *Hydrocyanic* or *Prussic acid.* This body is not an acid, although often called so. It is produced by the action of acids upon cyanides. It is an easily decomposed liquid of a rather pleasant odor, and intensely poisonous, a drop or two causing death almost instantaneously. It is used in medicine in a much diluted form. Hydrogen cyanide is produced in the fermentation of some glucosides.

Potassium cyanide, KCN, can be made by passing nitrogen over a mixture of potassium carbonate and charcoal.

$$K_2CO_3 + C_4 + N_2 = 2KCN + 3CO.$$

It is usually made from potassium ferrocyanide (*q. v.*). It is in white, fusible, deliquescent masses, and is a violent poison, probably because it easily furnishes HCN by the action of even feeble acids. It is extensively used in photography in order to dissolve the unaltered salts of silver, a double salt of potassium and silver resulting (KCN + AgCN). It is also employed in silver and gold plating.

The cyanides have a strong tendency to form double salts, and several of these are of much importance.

Potassium ferrous cyanide, FeCy$_2$ + 4KCy, usually called

potassium ferrocyanide and written K_4Cy_6Fe, is formed by heating a mixture of nitrogenous organic matter, iron scraps and potassium carbonate, and treating the mass with water. The salt forms large lemon-yellow crystals, which are not poisonous. It is much used in dyeing under the name of yellow prussiate of potash. Oxidizing agents convert the ferrous cyanide into ferric, and produce a body called

Potassium ferric cyanide, $Fe_2Cy_6 + 6KCy$, commonly called *potassium ferricyanide*, or red prussiate of potash. It forms large ruby-red crystals, soluble in water and used in dyeing. From these double cyanides others may be obtained, especially with iron. When ferrocyanide is added to a ferric salt, a blue precipitate of Prussian blue is formed; when ferricyanide is added to a ferrous salt, Turnbull's blue is formed.

Cyanogen forms several acids which have the same percentage composition, but different formulæ:

HCNO	Cyanic.
$H_2C_2N_2O_2$	Fulminic.
$H_3C_3N_3O_3$	Cyanuric.

This is called a *polymeric* series. We have also a sulphur acid, HCNS, which forms sulphocyanates.

AMINES AND AMIDES.

Most of the members of the nitrogen group combine with the hydrocarbon- and acid- radicles, producing substances which may be compared to amine, H_3N, or to ammonium compounds, and often having strong resemblance to these in general properties. They are, in fact, more conveniently studied by taking amine and ammonium chloride as types, and regarding the radicles present as substituting the hydro-

gen. They are named from the radicles present, with a prefix indicating the number of molecules, and with the name of the member of the nitrogen group. If referred to the type H_3N, they are called *amines;* if to the type NH_4Cl, *oniums.*

Amines.		Oniums.	
H_3N	Amine.	H_4NCl	Ammonium chloride.
$(C_2H_5)H_2N$	Ethylamine.	$(C_2H_5)H_3NCl$	Ethyl ammonium chloride.
$(C_2H_5)_2HN$	Diethylamine.	$(C_2H_5)_2(CH_3)HNCl$	Methyl diethyl ammonium chloride.
$(C_2H_5)_2(CH_3)N$	Methyl diethylamine.		
$(C_2H_5)_3P$	Triethyl phosphine.		
$(C_2H_5)_3B$	Triethyl borine.	$(C_2H_5)_4PI$	Tetrethyl phosphonium iodide.

It has also been found that bodies can be obtained with two, three and four atoms of nitrogen, phosphorus, etc., giving rise to diamines, triamines and tetramines.

$(C_2H_4)H_4N_2$ Ethene diamine $(C_2H_4)H_6N_2I_2$ Ethene diammonium iodide.

Arsenic forms with methyl a compound which is interesting from its remarkable properties, and because it was known for a long while before its composition was understood. This is arsen-dimethyl, $As(CH_3)_2$ (in the free state $As_2(CH_3)_4$). It is produced by distilling potassium acetate with arsenous anhydride. It was originally called *kakodyl*—a word derived from the Greek and referring to the disagreeable odor. It is highly poisonous, but yields an acid, $HAs(CH_3)_2O_2$, which is apparently non-poisonous, although containing over fifty per cent. As.

Several amines exist ready formed in plants. Among these are *Conia*, $C_8H_{15}N$, the active principle of water-hemlock, and *Nicotina*, $C_{10}H_{14}N_2$, the active principle of tobacco.

Natural Alkaloids. The medicinal virtues of plants are sometimes due to resins or volatile oils, but generally to nitrogenous bodies which are evidently related to the amines and oniums, but their exact nature is not made out. They are mostly insoluble in water, but soluble in alcohol. With acids they generally form crystalline salts, and are therefore called *alkaloids*. It will be sufficient to give a table showing the composition, sources and important properties of the more important:

Alkaloids.	Sources.	Properties.	Formulæ.
Conia	Water-hemlock	Colorless liquid	$C_8H_{15}N$
Nicotina	Tobacco	Oily liquid	$C_{10}H_{14}N_2$
Quinia	Peruvian bark	White crystals	$C_{20}H_{24}N_2O_2$
Cinchonia	"	" "	$C_{20}H_{24}N_2O$
Morphia	Opium	" "	$C_{17}H_{19}NO_3$
Atropia	Belladonna	" "	$C_{17}H_{23}NO_3$
Veratria	Hellebore	White powder	$C_{32}H_{52}N_2O_8$
Strychnia	Nux vomica	White crystals	$C_{21}H_{22}N_2O_2$
Caffeina	Coffee and tea	" "	$C_8H_{10}N_4O_2$
Urea, CH_4N_2O, is of animal origin.			

The fully-developed tissues and organs of animals are of such complicated composition that little information has been obtained in regard to them. They contain other elements besides those mentioned in the preceding compounds, sulphur, phosphorus and iron being the most important.

ATOMIC WEIGHTS, SYMBOLS AND ATOMICITIES.

Element.	Symbol.	Atomicity.	Atomic Weight.	Element.	Symbol.	Atomicity.	Atomic Weight.
Aluminum	Al	iv	27	Mercury	Hg	ʺ	200
Antimony	Sb	ʹʹʹ v	122	Molybdenum	Mo	vi	96
Arsenic	As	ʹʹʹ v	75	Nickel	Ni	ʺ	59
Barium	Ba	ʺ	137	Niobium	Nb	v	94
Beryllium	Be	ʺ	9.4	Nitrogen	N	ʹʹʹ v	14
Bismuth	Bi	ʹʹʹ v	210	Osmium	Os	vi	199
Boron	B	ʹʹʹ	11	Oxygen	O	ʺ	16
Bromine	Br	ʹ	80	Palladium	Pd	ʺ iv	106.5
Cadmium	Cd	ʺ	112	Phosphorus	P	ʹʹʹ v	31
Cæsium	Cs	ʹ	133	Platinum	Pt	iv	197.1
Calcium	Ca	ʺ	40	Potassium	K	ʹ	39.1
Carbon	C	iv	12	Rhodium	Rh	ʺ iv	104.3
Cerium	Ce	ʹʹʹ	138	Rubidium	Rb	ʹ	85.4
Chlorine	Cl	ʹ	35.4	Ruthenium	Ru	ʺ v	104.4
Chromium	Cr	iv	52.2	Selenium	Se	ʺ iv vi	79.5
Cobalt	Co	ʺ	58.8	Silicon	Si	iv	28
Copper	Cu	ʺ	63.4	Silver	Ag	ʹ	108
Didymium	Di	ʹʹʹ	144.8	Sodium	Na	ʹ	23
Erbium	E		169	Strontium	Sr	ʺ	87.5
Fluorine	F	ʹ	19	Sulphur	S	ʺ iv vi	32
Gallium	Ga		68.9	Tantalum	Ta	iv	182
Gold	Au	ʹʹʹ v	196.7	Tellurium	Te	ʺ iv vi	129
Hydrogen	H	ʹ	1	Thallium	Tl	ʹ ʹʹʹ	204
Indium	In	ʹʹʹ	113.4	Thorium	Th		231.5
Iodine	I	ʹ	127	Tin	Sn	iv	118
Iridium	Ir	ʺ iv	198	Titanium	Ti	iv	50
Iron	Fe	ʺ iv	56	Tungsten	W	iv	184
Lanthanum	La	ʹʹʹ	139	Uranium	U	ʺ	120
Lead	Pb	ʺ	207	Vanadium	V	ʹʹʹ v	51.3
Lithium	Li	ʹ	7	Yttrium	Y	ʺ	92
Magnesium	Mg	ʺ	24	Zinc	Zn	ʺ	65
Manganese	Mn	ʺ iv	55	Zirconium	Zr	iv	89.5

THE METRIC SYSTEM.

The **Metric System** is a decimal system of weights and measures, and derives its name from the **meter** which is the unit of the system.

Note.—The metric system originated in France, and was finally made obligatory in that country in 1841. In 1866 its use in the United States was authorized by Act of Congress.

The length of the meter was intended to be one ten-millionth of the distance from the equator to either pole, measured at the level of the sea; but it is in reality a trifle less.

The higher denominations of any measure, obtained by multiplication of the unit, are named by prefixing to the name of the unit of that measure, the *Greek* numerals Deka, 10, Hecto, 100, Kilo, 1000, or Myria, 10000.

The lower denominations of any measure, obtained by division of the unit, are named by prefixing to the name of the unit of that measure, the *Latin* numerals Deci, $\frac{1}{10}$, Centi, $\frac{1}{100}$, or Milli, $\frac{1}{1000}$.

ABBREVIATIONS.

The name of each leading unit is abbreviated by writing its first letter after the number denoting the given quantity; 1 l. = 1 liter, 5 m. = 5 meters, etc. The Greek names of the higher orders of units are abbreviated in capitals; Hg. = Hektogram, Kg. = Kilogram, etc. The Latin names of the lower orders of units are abbreviated in small letters; ds. = decistere, mm. = millimeter, etc. In square measure and cubic measure, sq. and cu. = square and cubic respectively; 5 sq. Dm. = 5 square Dekameters; 9 cu. dm. = 9 cubic decimeters.

MEASURES OF LENGTH.

The unit of length is the **Meter.**

10 Millimeters = 1 Centimeter = .3937079 inches.
10 Centimeters = 1 Decimeter   = 3.937079   "
10 Decimeters   = 1 **METER**   = 39.37079   "
10 Meters   = 1 Dekameter   = 393.7079   "
10 Dekameters   = 1 Hektometer = 3937.079   "
10 Hektometers = 1 Kilometer   = 39370.79   "
10 Kilometers   = 1 Myriameter = 393707.9   "

Note.—The Meter is used in the same manner as our **yard, and the Kilometer** as our mile.

MEASURES OF CAPACITY.

		DRY MEASURE.	LIQUID MEASURE.
10 Milliliters	=1 Centiliter	=0.6102 cu. in.	or 0.338 fl.oz.
10 Centiliters	=1 Deciliter	=6.1022 "	" 0.845 gi.
10 Deciliters	=1 **LITER**	= 0.908 qt.	" 1.0567 qt.
10 Liters	=1 Dekaliter	= 9.08 "	" 2.6417 gal.
10 Dekaliters	=1 Hektoliter	=2.8375 bu.	"26.417 "
10 Hektoliters	=1 Kiloliter	=28.375 "	"264.17 "
10 Kiloliters	=1 Myrialiter	=283.75 "	"2641.7 "

MEASURES OF WEIGHT.

The unit of weight is the **Gram.**

10 Milligrams = 1 Centigram = 0.1543 grains.
10 Centigrams = 1 Decigram   = 1.5432   "
10 Decigrams   = 1 **GRAM**   = 15.432   "
10 Grams   = 1 Dekagram   = 0.3527 oz. Avoir.
10 Dekagrams   = 1 Hektogram = 3.527 "   "
10 Hektograms = 1 Kilogram   = 2.2046 lb.   "
10 Kilograms   = 1 Myriagram = 22.046 "   "
10 Myriagrams = 1 Quintal   = 220.46 "   "
10 Quintals   = 1 Tonneau   = 2204.6 "   "

CUBIC MEASURE.

CUBIC INCHES.

1 Cu. In. = 16.39 Cu. Centimeters	6 Cu. In. = 98.32 Cu. Centimeters.
2 " = 32.77 "	7 " = 114.70 "
3 " = 49.16 "	8 " = 131.09 "
4 " = 65.54 "	9 " = 147.47 "
5 " = 81.93 "	10 " = 163.86 "

CUBIC FEET.

1 Cu. Ft. = 28.32 Cu. Decimeters.	6 Cu. Ft. = 169.89 Cu. Decimeters.
2 " = 56.63 "	7 " = 198.20 "
3 " = 84.95 "	8 " = 226.52 "
4 " = 113.26 "	9 " = 254.84 "
5 " = 141.58 "	10 " = 283.15 "

CUBIC YARDS.

1 Cu. Yd. = .7645 Cu. Meters.	6 Cu. Yd. = 4.5870 Cu. Meters.
2 " = 1.5290 "	7 " = 5.3515 "
3 " = 2.2935 "	8 " = 6.1160 "
4 " = 3.0580 "	9 " = 6.8806 "
5 " = 3.8225 "	10 " = 7.6451 "

LIQUID MEASURE.

GILLS.

1 Gill = 1.1831 Deciliters.	3 Gills = 3.5493 Deciliters.
2 " = 2.3662 "	4 " = 4.7325 "

PINTS.

1 Pint = 4.7325 Deciliters.	2 Pints = 9.4650 Deciliters.

QUARTS.

1 Quart = 9.4650 Deciliters.	3 Quarts = 2.8395 Liters.
2 " = 1.8930 Liters.	4 " = 3.7860 "

GALLONS.

1 Gallon = 3.786 Liters.	6 Gallons = 2.271 Dekaliters.
2 " = 7.572 "	7 " = 2.650 "
3 " = 1.135 Dekaliters.	8 " = 3.028 "
4 " = 1.514 "	9 " = 3.407 "
5 " = 1.893 "	10 " = 3.786 "

DRY MEASURE.

PINTS.

1 Pint = 5.5067 Deciliters.   2 Pints = 1.1013 Liters.

QUARTS.

1 Quart	= 1.1013 Liters.		5 Quarts	= 5.5067 Liters.	
2 "	= 2.2027	"	6 "	= 6.6081	"
3 "	= 3.3040	"	7 "	= 7.7094	"
4 "	= 4.4054	"	8 "	= 8.8108	"

PECKS.

1 Peck = 8.8108 Liters.   3 Pecks = 2.6432 Dekaliters.
2 " = 1.7621 Dekaliters.   4 " = 3.5243   "

BUSHELS.

1 Bushel	= 3.524 Dekaliters.		6 Bushels	= 2.114 Hektoliters.	
2 "	= 7.048	"	7 "	= 2.467	"
3 "	= 1.057 Hektoliters.		8 "	= 2.819	"
4 "	= 1.409	"	9 "	= 3.171	"
5 "	= 1.762	"	10 "	= 3.524	"

AVOIRDUPOIS WEIGHT.

OUNCES.

1 Ounce	= 2.835 Dekagrams.		9 Ounces	= 2.551 Hektograms.	
2 "	= 5.671	"	10 "	= 2.835	"
3 "	= 8.506	"	11 "	= 3.119	"
4 "	= 1.134 Hektograms.		12 "	= 3.402	"
5 "	= 1.417	"	13 "	= 3.685	"
6 "	= 1.701	"	14 "	= 3.969	"
7 "	= 1.984	"	15 "	= 4.252	"
8 "	= 2.268	"	16 "	= 4.536	"

POUNDS.

1 Pound	= 4.5359 Hektograms.		6 Pounds	= 2.7216 Kilograms.	
2 "	= 9.0718	"	7 "	= 3.1751	"
3 "	= 1.3608 Kilograms.		8 "	= 3.6227	"
4 "	= 1.8144	"	9 "	= 4.0823	"
5 "	= 2.2680	"	10 "	= 4.5359	"
2000 "	= .9072 Tonneau.		2240 "	= 1.0160 Tonneaux.	

APOTHECARIES' WEIGHT.

GRAINS.

1 Grain	= 6.480 Centigrams.		11 Grains	= 7.1280 Decigrams.	
2 "	= 1.296 Decigrams.		12 "	= 7.7760	"
3 "	= 1.944	"	13 "	= 8.4240	"
4 "	= 2.592	"	14 "	= 9.0720	"
5 "	= 3.240	"	15 "	= 9.7200	"
6 "	= 3.888	"	16 "	= 1.0368 Grams.	
7 "	= 4.536	"	17 "	= 1.1016	"
8 "	= 5.184	"	18 "	= 1.1664	"
9 "	= 5.832	"	19 "	= 1.2312	"
10 "	= 6.480	"	20 "	= 1.2960	"

SCRUPLES.

2 Scruples = 2.5920 Grams.   3 Scruples = 3.8879 Grams.

DRACHMS.

1 Drachm	= 3.8879 Grams		5 Drachms	= 1.9440 Dekagrams.	
2 "	= 7.7758	"	6 "	= 2.3328	"
3 "	= 1.1664 Dekagrams.		7 "	= 2.7216	"
4 "	= 1.5552	"	8 "	= 3.1103	"

OUNCES.

1 Ounce	= 3.1103 Dekagrams.		7 Ounces	= 2.1772 Hektograms.	
2 "	= 6.2206	"	8 "	= 2.4883	"
3 "	= 9.3309	"	9 "	= 2.7993	"
4 "	= 1.2441 Hektograms.		10 "	= 3.1103	"
5 "	= 1.5552	"	11 "	= 3.4214	"
6 "	= 1.8662	"	12 "	= 3.7324	"

POUNDS.

1 Pound	= 3.7324 Hektograms.		6 Pounds	= 2.2394 Kilograms.	
2 "	= 7.4648	"	7 "	= 2.6127	"
3 "	= 1.1197 Kilograms.		8 "	= 2.9859	"
4 "	= 1.4929	"	9 "	= 3.3592	"
5 "	= 1.8662	"	10 "	= 3.7324	"

Note.—The Troy pound, ounce, and grain have the same values as in Apothecaries' Weight.

APOTHECARIES' LIQUID MEASURE.

MINIMS.

1 Minim = .0616 Milliliters.	6 Minims = .3697 Milliliters.
2 " = .1232 "	7 " = .4313 "
3 " = .1848 "	8 " = .4930 "
4 " = .2465 "	9 " = .5546 "
5 " = .3081 "	10 " = .6162 "

FLUID DRACHMS.

1 Fl. Dr. = 3.6973 Milliliters.	5 Fl. Dr. = 1.8486 Centiliters.
2 " = 7.3946 "	6 " = 2.2184 "
3 " = 1.1092 Centiliters.	7 " = 2.5881 "
4 " = 1.4789 "	8 " = 2.9578 "

FLUID OUNCES.

1 Fl. Oz. = 2.9578 Centiliters.	9 Fl. Oz. = 2.6620 Deciliters.
2 " = 5.9156 "	10 " = 2.9578 "
3 " = 8.8734 "	11 " = 3.2536 "
4 " = 1.1831 Deciliters.	12 " = 3.5494 "
5 " = 1.4789 "	13 " = 3.8452 "
6 " = 1.7747 "	14 " = 4.1409 "
7 " = 2.0705 "	15 " = 4.4367 "
8 " = 2.3662 "	16 " = 4.7325 "

PINTS.

1 Pint = 4.7325 Deciliters.	5 Pints = 2.3662 Liters.
2 " = 9.4650 "	6 " = 2.8395 "
3 " = 1.4197 Liters.	7 " = 3.3127 "
4 " = 1.8930 "	8 " = 3.7860 "

GALLONS.

1 Gallon = 3.786 Liters.	6 Gallons = 2.271 Dekaliters.
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INDEX.

BUTLER'S

Elements

of

Chemistry

Leffmann